

Inner Space: The Structure of the Atom

by Samuel Glasstone

WORLD OF THE ATOM SERIES

ATOM

Although the atomic substratum is not immediately apparent to our senses, its activities exert their effects upon many phenomena that can be observed. They influence the course of events in the remote stars of the heavens, in our earthly environment, in all living creatures, including ourselves.

Victor Guillemin



The Author

Samuel Glasstone—Ph.D. (1922), D.Sc. (1926) University of London, in Physical Chemistry—holds a pre-eminent position as a lucid expositor of scientific subject matter. He has written 34 books, sometimes with the co-operation of other scientists. In 1959 the American Society of Mechanical Engineers awarded him the Worcester Reed Warner Medal in recognition of his "outstanding contribution to permanent engineering literature in (his) writings on atomic energy". In 1968 he received the Arthur Holly Compton Award from the American Nuclear Society for "his distinguished contributions to nuclear science and engineering education".



Dr. Glasstone's best known book in the nuclear field is *Sourcebook on Atomic Energy*; first published in 1950 and revised in 1958 and 1967, it is still a best seller. He has also written for scientists and engineers about reactor theory, nuclear engineering, nuclear weapons, and controlled thermonuclear research, as well as various aspects of physical chemistry.

For the World of the Atom Series, he wrote *Controlled Nuclear Fusion* and coauthored *Atomic Energy and Your World*.

The Cover

"Hanging Mobile", Alexander Calder, 1936. The Museum of Modern Art, New York, Mary Callery Collection.

The World of the Atom Series

Nuclear energy is playing a vital role in the life of every man, woman, and child in the United States today. In the years ahead it will affect increasingly all the peoples of the earth. It is essential that all Americans gain an understanding of this vital force if they are to discharge thoughtfully their responsibilities as citizens and if they are to realize fully the myriad benefits that nuclear energy offers them.

The United States Atomic Energy Commission provides this booklet to help you achieve such understanding.

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Inner Space: The Structure of the Atom

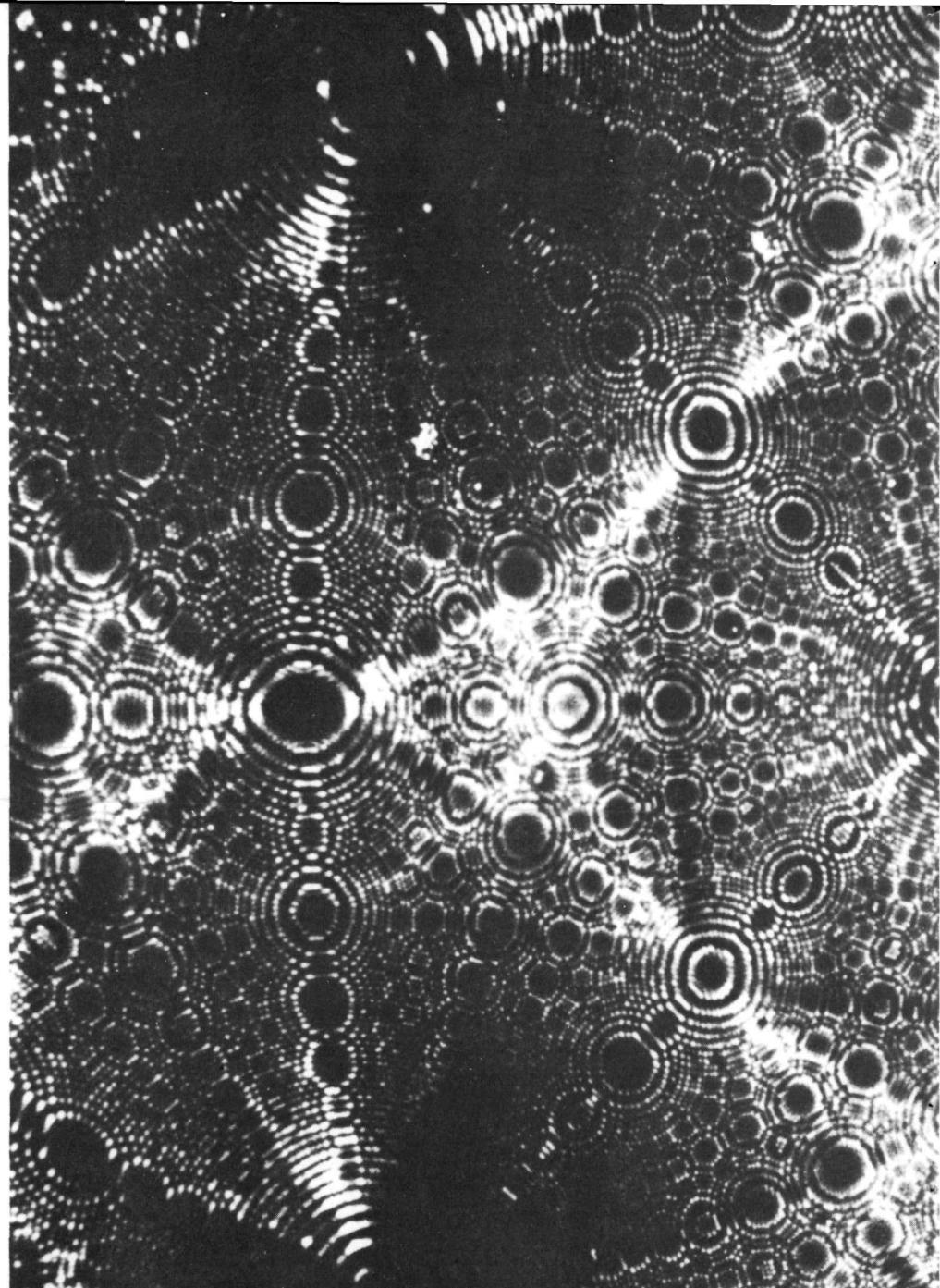
by **Samuel Glasstone**

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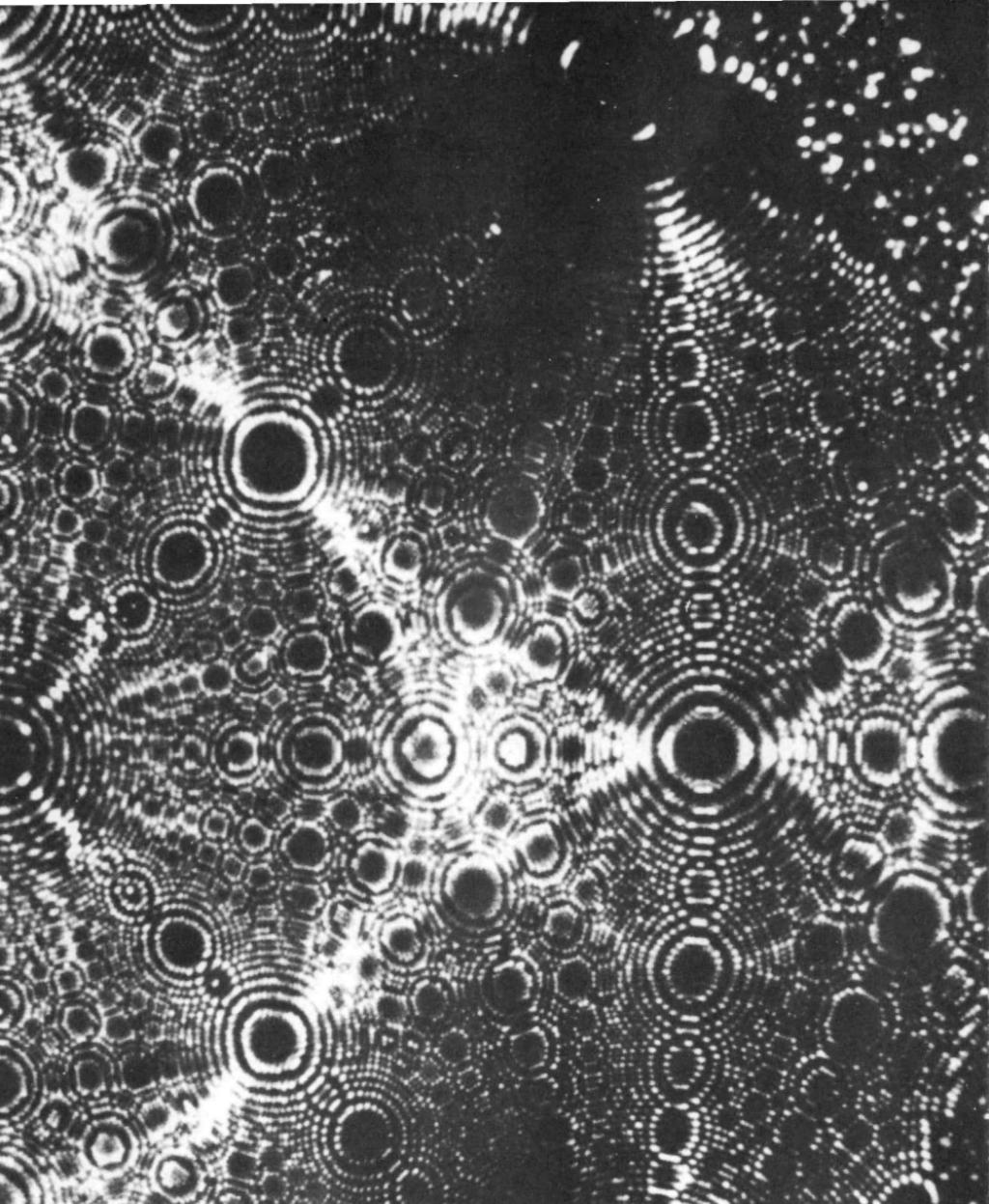
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Field-ion micrograph of a crystal. Each tiny dot is a single atom.



Matter, though divisible to an extreme degree, is nevertheless not infinitely divisible . . . The existence of ultimate particles of matter can scarcely be doubted though they are probably much too small ever to be exhibited by microscopic improvements.

John Dalton

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INTRODUCTION

Elements and Atoms

In considering the makeup of the universe, Greek philosophers, over 2000 years ago, thought that all things were composed of four "elements"—earth, air, fire, and water. This idea was widely accepted until the latter part of the 18th century when discoveries made by scientists began to throw doubts on its validity. For example, in 1774 the French chemist Antoine L. Lavoisier showed that air was not an elementary (or single) substance, but was a mixture of two gases, oxygen and nitrogen. A few years later, in 1789, the work of Henry Cavendish in England proved water to be compounded of hydrogen and oxygen. Thus, two of the so-called "elements" of the ancients, air and water, were not elements at all.

It is to Lavoisier that we owe the modern concept of an element as a substance containing only one kind of matter. Everything is made up of these elements. Some substances, such as carbon, iron, and gold, are themselves elements. Some are tight combinations of elements, called chemical compounds; examples are water (hydrogen and oxygen), salt (sodium and chlorine), and sugar (carbon, hydrogen, and oxygen). Other substances such as air, which is mainly



Portraits de M^r & M^{me} Lavoisier

Antoine Lavoisier and his wife. Engraving made from a painting by Jacques Louis David.

oxygen and nitrogen, are mixtures of elements, while still others, such as a salt or sugar in solution, are mixtures of compounds. At the present time, 90 elements are known to exist on the earth, and some 15 others, which are unstable (radioactive), have been produced in the laboratory.*

For many years scientists asked themselves: What makes one element different from another? In order to answer this question we must consider a further aspect of the structure of matter.

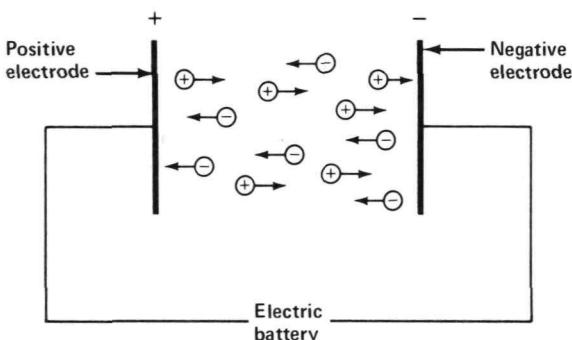
In addition to giving us the first concept of an element, the ancient Greeks also developed the idea of *atoms* as the smallest, indivisible units of which all material things are composed. The atom is now regarded as the smallest possible particle of an element that retains the identity of that element. It will be seen in due course that atoms can be broken into still smaller parts, but these parts do not have the properties of the element. The atoms of an element determine the characteristics of that particular element. All the atoms of a given element are essentially alike, but they differ from the atoms of any other element. One of the purposes of this booklet is to explain how the atoms of various elements differ from one another.

The Unit of Electrical Charge

Under certain conditions an atom or a chemical combination of atoms, called a molecule, can acquire either a positive or a negative charge of electricity. Such electrically charged atoms or molecules are known as *ions*. Ions can be present in a gas, in a liquid (solution), or in a solid. In gases and solutions, and occasionally in solids, the ions are free to move and their motion can be directed by the application of an

*For information about some of these unstable elements, see *Synthetic Transuranium Elements*, another booklet in the World of the Atom series.

electrical field (or voltage). The word ion, meaning "traveler" in Greek, owes its origin to this directed motion of the electrically charged particles. In an electrical field positive ions travel in one direction, from the positive pole (or electrode) toward the negative pole (or electrode), whereas negative ions move in the opposite direction (see the figure). In fact, it is this motion of charged particles in opposite directions that is responsible for the passage of electricity through gases and solutions.



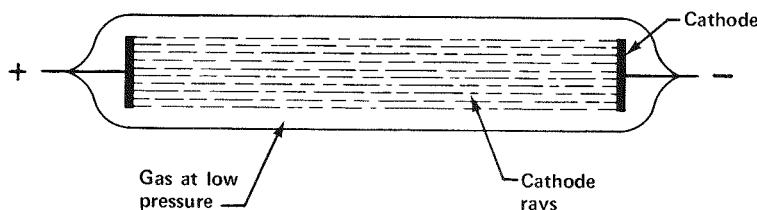
In an electric field, positive ions move toward the negative electrode and at the same time the negative ions move toward the positive electrode.

Physicists reached an important conclusion from studies of the behavior of ions, especially from work on the passage of electricity through solutions carried out during the 1830s by Michael Faraday in England. If elements are composed of atoms then there must also be an "atomic" unit of electricity. That is to say, an electrically charged atom or molecule (an ion) must always carry a whole (integral) number of units of electric charge. An ion may have one, two, three, or more units of electrical charge, but never a fractional charge. To this fundamental unit of electricity the Irish physicist G. Johnstone Stoney in 1874 gave the name *electron*. It will be seen shortly that this word is now used in a different although related sense.



The Universal Electron

When a discharge of electricity is passed between two electrodes in a gas at low pressure (as shown in the figure), various luminous phenomena are observed, of which neon signs and fluorescent lamps provide familiar examples.



When electricity is passed through a gas at low pressure, various luminous phenomena are observed, including the production of cathode rays

Among the observations made by early experimenters using electrical discharges in gases is that the cathode (the electrode connected to the negative side of the electrical source) emits what are called *cathode rays*. The nature of these rays was the subject of much argument during the later years of the 19th century. Then, in 1897, the English physicist J. J. Thomson, whose work had a profound effect, both direct and indirect, on our understanding of the atom, finally solved the problem. He proved that cathode rays consist of particles carrying a negative electrical charge. These particles were always the same, regardless of the cathode material or the nature of the gas through which the electrical discharge was passed.

In subsequent years, Thomson showed that particles with essentially the same properties as those in cathode rays were produced in other ways. For example, such particles are

Michael Faraday

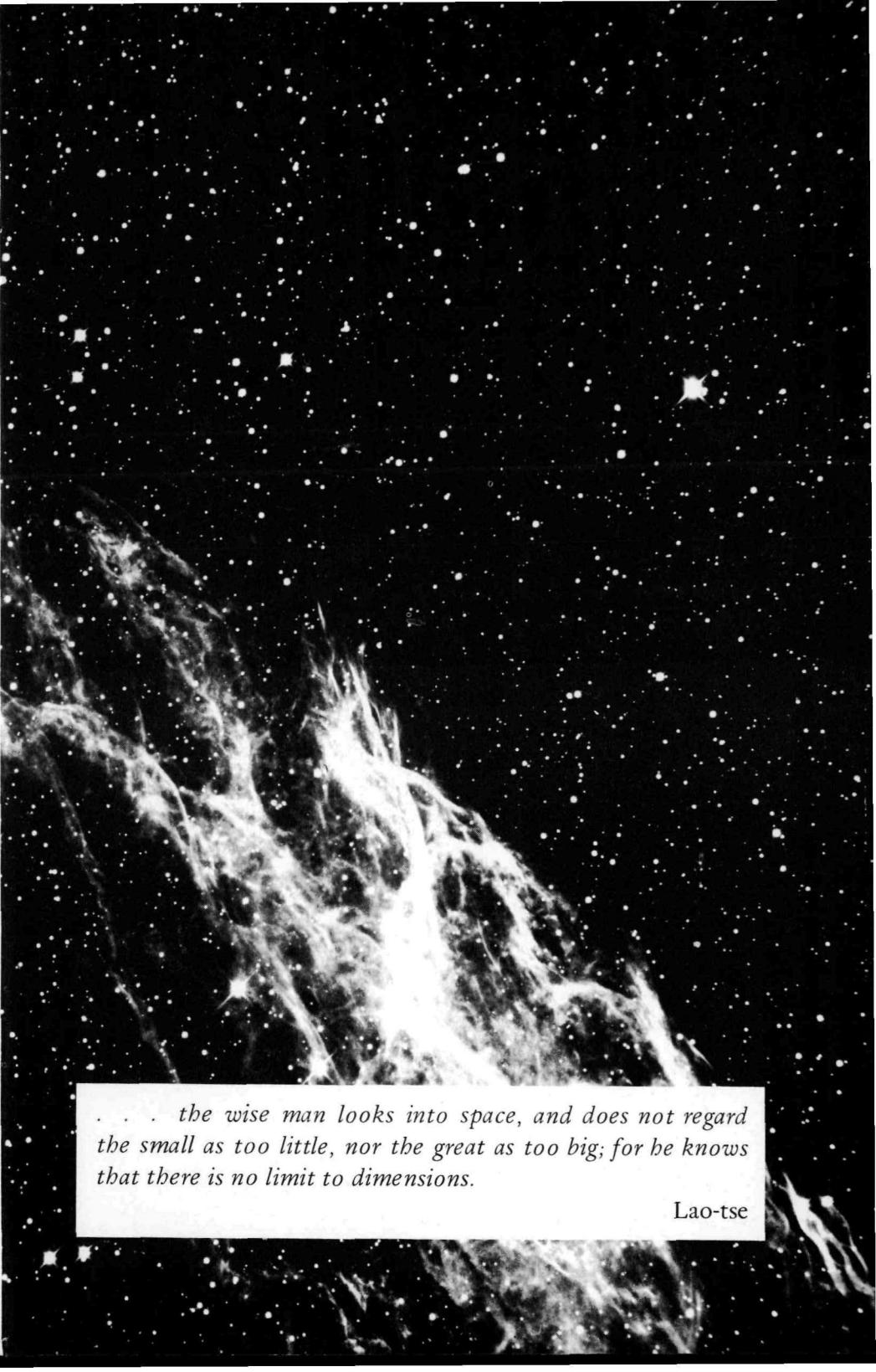


J. J. Thomson in his laboratory. On the right are early X-ray pictures.

generated when ultraviolet light falls on certain metals (photoelectric effect) and they are emitted from a hot filament of a metal or carbon (thermionic effect). These particles or *corpuscles*, as Thomson called them, were all found to carry negative electrical charges of the same magnitude.

Because the charge carried by the corpuscles was identical with the unit of electricity carried by ions, either gaseous or in solution, the name "electron" became associated with the particles themselves, although it was originally intended for the magnitude of the unit charge. Thus, the particles, present in cathode rays and formed by the photoelectric and thermionic effects, are now known as electrons. The unit electrical charge they carry is referred to as the electronic charge.





. . . the wise man looks into space, and does not regard the small as too little, nor the great as too big; for he knows that there is no limit to dimensions.

Lao-tse

THEORIES OF ATOMIC STRUCTURE

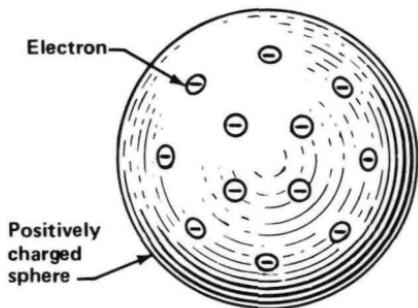
Thomson's Theory of the Atom

The fact that the same particles—electrons—could be produced from different substances in different ways led Thomson in 1898 to an important idea. He suggested that all atoms contain electrons but in different numbers. Since the atom as a whole is electrically neutral, that is to say, it has no net electrical charge, it must contain a positive charge to neutralize (or balance) the negative charge of the electrons. Thomson thought that this positive charge was distributed throughout the volume of the atom in what he later called “a sphere of uniform positive electrification”, as indicated in the figure.

Apart from the somewhat vague nature of this sphere of positive electrification, there was another difficulty with Thomson's theory of the atom which he himself realized. The mass of an electron is very much less than that of an atom. In fact, a hydrogen atom, the lightest atom known, is almost two thousand times as heavy as an electron.* Other atoms are still heavier in comparison with an electron. Consequently, if the electrons were wholly (or largely) responsible for the mass of an atom, then a single atom, especially of the heavier elements, would have to contain many thousands of electrons.

Experiments on the behavior of light and X rays in gases, however, indicated that this could not be so, and that the number of electrons in an atom must be fairly small. It seemed, therefore, that the positive charge must make a significant contribution to the mass of the atom. This idea turned out to be substantially correct; as will be seen shortly, nearly all the mass of the atom is contributed by the carriers

*More precisely, a hydrogen atom has the same mass as 1837 electrons.



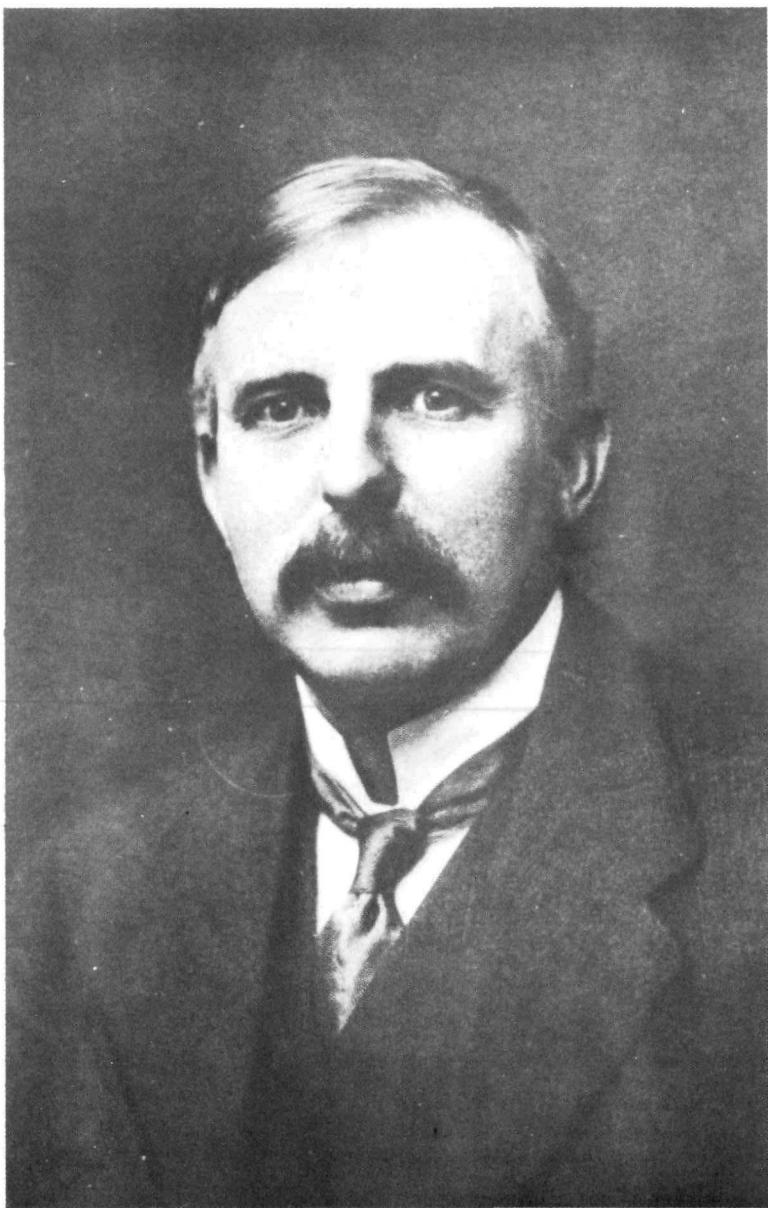
The model of an atom as proposed by J. J. Thomson. The positive charge uniformly distributed throughout the sphere would balance the negative charge of the electrons.

of the positive charge. But at the time, it was difficult to reconcile this fact with the supposed nature of the positive charge distribution. Although he was thus baffled in the development of a satisfactory theory of the composition, or structure, of atoms, Thomson opened up a new era in the study of the atom by calling attention to the universality of the electron and by showing that a neutral atom might actually be a system of negative (electrons) and positive charges.

Scattering of Alpha Particles

The modern views concerning the structure of the atom arose quite unexpectedly from experiments on the behavior of *alpha particles* emitted by some radioactive substances.* These particles are relatively heavy—about 7350 times the mass of an electron—and carry two positive charges of electricity. In 1906, Ernest Rutherford—at that time in Canada—had noted that when a narrow beam of alpha particles fell upon a photographic plate after passing through

*For short reviews of radioactivity, see the booklets *Accelerators* and *Our Atomic World* in this series.

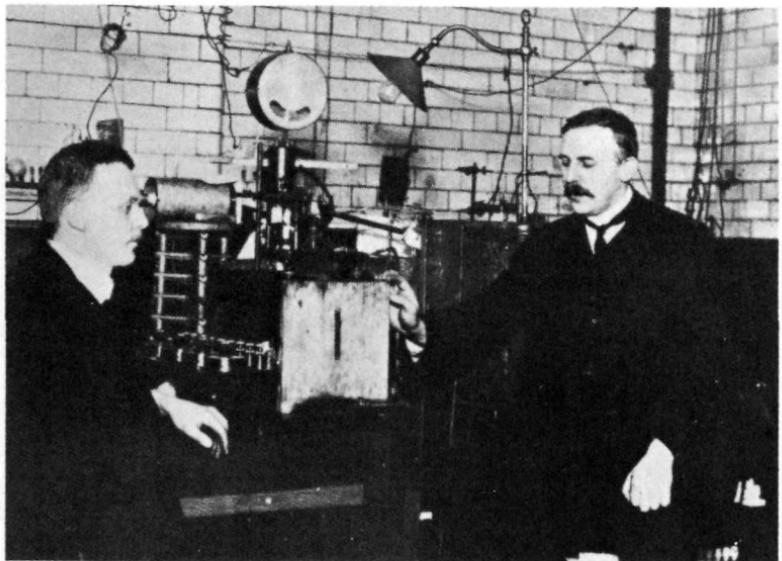


Ernest Rutherford

a very thin metal sheet (or foil), the resulting trace was not sharp but faded off at the edges. On the other hand, alpha particles which did not pass through the metal foil gave a trace with sharp edges.

The alpha particles were apparently deflected (scattered) from their straight-line course by the foil.

In further studies of alpha-particle scattering, Rutherford's assistants, Hans Geiger and Ernest Marsden, in Manchester, England, noted that the majority of the alpha particles continued in their original direction of motion or were scattered only to a slight extent upon passing through a metal foil. However, to their great surprise, the scientists observed that some of the alpha particles were deflected through large angles. A few of the particles even emerged from the same side of the foil as they had entered (see the figure on the next page). "It was about as credible," said Rutherford, "as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

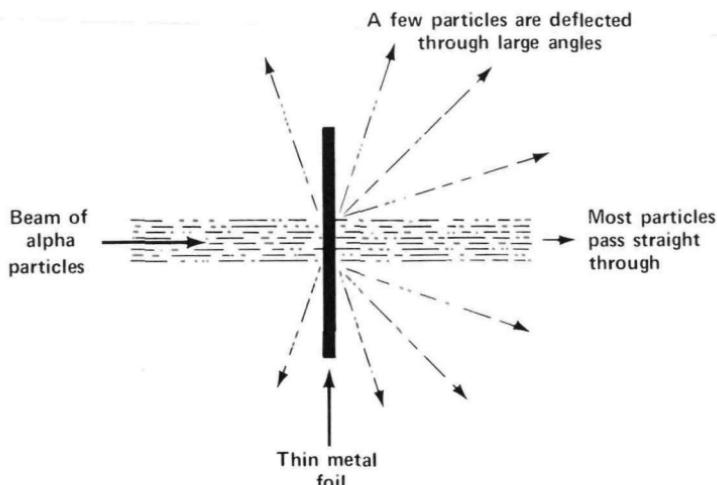


Hans Geiger (left) and Ernest Rutherford at Manchester University about 1910.

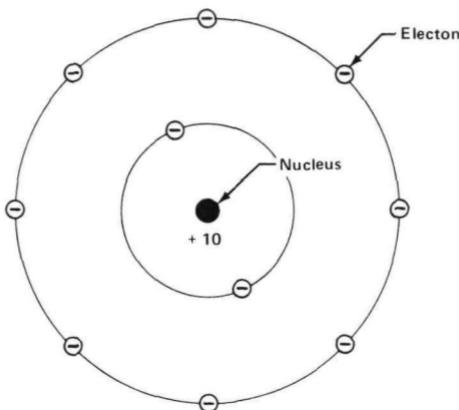
The Nuclear Theory of the Atom

At first Rutherford attempted to explain the unexpected large-angle scattering of the alpha particles in terms of a succession of deflections through many small angles, all in the same direction. Such small-angle deflections might perhaps result, he thought, from interactions of the alpha particles with the electrons in the foil atoms. Later, Rutherford realized that this was not a satisfactory explanation of the occasional scattering of alpha particles through large angles. Further consideration then led him, in 1911, to account for the scattering on the basis of a new concept of the structure of the atom. This theory of atomic structure has become so well established that no other is now given serious consideration.

Rutherford argued that Thomson's idea of a uniform distribution of positive charge throughout an atom is wrong. Actually, he said, the positive charge (and most of the mass) is concentrated in a very small region, which he later called



When a narrow beam of alpha particles passes through a thin metal foil, most of the particles continue in the original direction of motion but a few are deflected (scattered) through large angles.



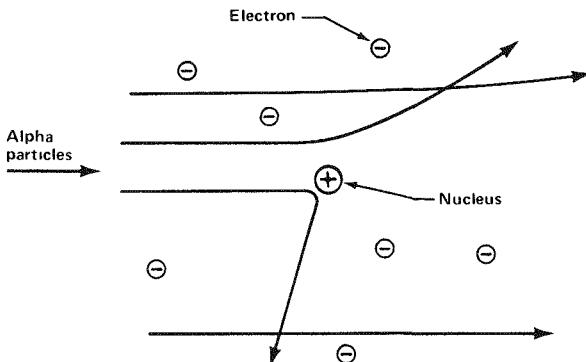
Representation of an atom (of neon) with a nucleus carrying 10 positive charges surrounded by 10 electrons. (If this were drawn to scale, the nucleus would be about 10^{-5} (a hundred-thousandth) inch in diameter and the electrons would be somewhat smaller.)

the *nucleus*, located at the center of the atom. The negatively charged electrons, relatively small in number, occupy the space surrounding the nucleus (see the figure above). Because the electrons are small in comparison with the size of the atom, much of the interior of the atom is empty.*

The nuclear atom of Rutherford provided an entirely satisfactory explanation of the scattering of alpha particles. Since much of the interior of an atom is empty space, the great majority of alpha particles striking a very thin sheet of metal will pass right through it. Some scattering is caused by interaction with electrons, but this is small. A few alpha particles, however, will come fairly close to the nuclei at the center of the atoms. Then the positively charged alpha particle will be strongly repelled by the concentrated positive charge on the small nucleus, and large-angle scattering will result (see the figure on the next page).

In order to verify this picture of alpha-particle scattering, and of the nuclear atom, Rutherford investigated the

*The reason why matter is often solid (or liquid) in spite of the "emptiness" of the atom is that the electrons usually behave as if they occupied the whole volume of the atom.



The alpha particles that come close to the nucleus are scattered through large angles.

situation mathematically and derived an equation which could be tested by experiment. This equation related the numbers of particles scattered at various angles to the magnitude of the positive charge on the atomic nucleus, the thickness of the scattering foil, and the initial velocity (or energy) of the alpha particles. The results of extensive measurements made by Geiger and Marsden with seven different scattering materials and with alpha particles of different energies were found to be in good agreement with Rutherford's equation. The experiments thus provided strong evidence for the underlying assumption that an atom has at its center a positively charged nucleus which is small in comparison with the atom as a whole.

The Nuclear Charge and Mass

Two questions may be asked at this point: (1) What is the size (or radius) of a nucleus in comparison with that of the atom? (2) What is the magnitude of the positive charge on the nucleus? The latter question is of particular importance because of its implications. For example, it has a bearing on the nuclear mass. Since the atom is electrically neutral, the

number of electrons that occupy the space around the nucleus, called the *orbital electrons*, must be equal to the number of positive charges on the nucleus. Consequently, the magnitude of this positive charge will tell us how many orbital electrons there are in the given atom.

Furthermore, since the mass of a single electron is known, it is now possible to determine the mass of all the orbital electrons in the atom. The total mass of any atom can be found either from chemical determinations of the atomic weight or by direct measurement in an instrument called a *mass spectrometer*. By subtracting the mass of the electrons from the mass of the atom, the mass of the nucleus can be obtained.

The determination of the size of the nucleus requires some knowledge of the charge, and so we will consider this point first. Although the early experiments on the scattering of alpha particles were not of high accuracy, they did give a rough indication of the number of unit (electronic) positive charges carried by a nucleus. It appeared that for any element this number is, very approximately, half the atomic weight of the element.* This result was in general agreement with observations on the scattering of X rays, which depends on the electrons in the atom. Charles G. Barkla, in 1911, concluded from X-ray scattering that for several light elements the number of electrons was about half the atomic weight. The number of electrons in an atom should, of course, be equal to the number of unit positive charges on the nucleus.

Exact values of the nuclear charge will be given later, but for the time being we will assume it to be a whole number that is about half the atomic weight. Hence, if the atomic weight is represented by the symbol A , the positive charge on

*An adequate definition of atomic weight, for the present purpose, is the mass of the atom of any element compared with the mass of an atom of hydrogen, the lightest element.

the nucleus, which is also the number of orbital electrons in the atom, is roughly $0.5A$. The mass of an electron is about one two-thousandth (0.0005) of the mass of a hydrogen atom. Since we are taking the latter mass to be the unit of atomic weight, the mass of an electron is about 0.0005. The total mass of the $0.5A$ electrons in any atom is thus $0.0005 \times 0.5A$, which equals $0.00025A$ on the atomic weight scale. We thus have the following (approximate) result

$$\text{Total mass of the atom} = A$$

$$\text{Mass of the electrons} = 0.00025A$$

Hence, by subtraction,

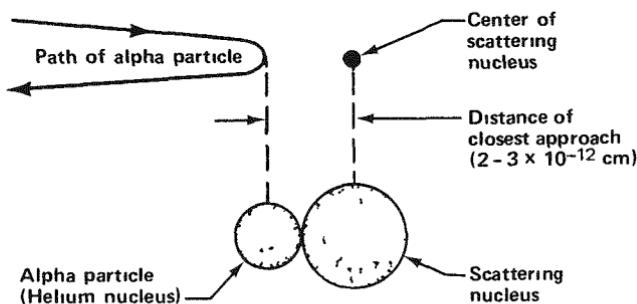
$$\text{Mass of the nucleus} = 0.99975A$$

In other words, the nucleus carries about 99.975% of the total mass of the atom, as well as all its positive charge. By using exact values of the number of positive charges on a nucleus, it has been found that, for atoms of different elements, 99.94 to 99.98% of the mass of the atom is located in the nucleus. We can see, therefore, that the nucleus carries almost all the mass of the atom, the electrons carry only a small fraction of 1%.

The Nuclear Size

According to Rutherford's calculations, the greater the angle through which an alpha particle is scattered, the more closely has it approached the atomic nucleus before being turned aside. By determining the maximum scattering angle for alpha particles of known energy, it is possible to calculate the distance of closest approach between the alpha particle and the nucleus of the scattering material, provided the charge on the nucleus is known. In this manner, Rutherford found that the closest distance to which an alpha particle could approach the nucleus of a heavy element, such as silver or gold, was a few (2 or 3) times 10^{-12} cm. This distance of closest approach may be taken as the sum of the radii of the

nucleus and an alpha particle (see the figure below) The alpha particle is known to be the nucleus of a helium atom Hence, it follows that the radius of a nucleus is of the order of 10^{-12} cm As we will see shortly, later measurements gave values between 10^{-13} and 10^{-12} cm



The distance of closest approach of an alpha particle to a scattering nucleus occurs when the alpha particle's direction of motion is reversed The distance is taken as the sum of the radii of the alpha particle (helium nucleus) and the scattering nucleus

In recent years, more accurate experimental methods have been developed for finding the radius of a nucleus Several of these depend on the scattering of charged particles, especially of electrons of very high energy The results of such measurements may be summarized, with a reasonable degree of accuracy for all but the lightest elements, by the expression

$$R = 1.25 \times 10^{-13} \sqrt[3]{A} \text{ cm}$$

where R is the radius of the nucleus in centimeters and A is the mass number of the atom the whole number closest to the atomic weight (see p 35) Thus, the radius of a nucleus is roughly proportional to the cube root of its mass, since most of the mass of the atom is in the nucleus

For the common atom of carbon, for example, A is 12 and $\sqrt[3]{12}$ is 2.29; hence the radius of the nucleus is about 2.9×10^{-13} cm. The heaviest atom in nature is uranium with $A = 238$; since $\sqrt[3]{238} = 6.19$, its nuclear radius is almost 7.8×10^{-13} cm. The unit of 10^{-13} cm, which is convenient for expressing the radius of a nucleus and of nuclear distances in general, is called 1 fermi, in honor of Enrico Fermi, an Italian-American physicist, who made outstanding contributions to the physics of the atom. Hence, we can say that nuclear radii lie within the range of about 2 to 8 fermis, with the value increasing according to the cube root of the mass of the atom.

We are now in a position to compare the size of the nucleus with that of the whole atom. Although the lighter atoms may have smaller radii and the more massive ones larger radii, a good average value for the radius of an atom is 2×10^{-8} cm. For comparison we may take about 5×10^{-13} cm to be the average radius of a nucleus. Thus the radius of the whole atom is $(2 \times 10^{-8})/(5 \times 10^{-13}) = 4 \times 10^4$ (or 40,000) times as great as that of the nucleus. This means that if a nucleus were magnified to the size of an orange, the atom would be almost 2 miles across! Such a comparison should make it clear that the atom has a very "empty" structure. It is true that the orbital electrons occupy the region around the nucleus, but the electrons themselves are quite small and take up little space.

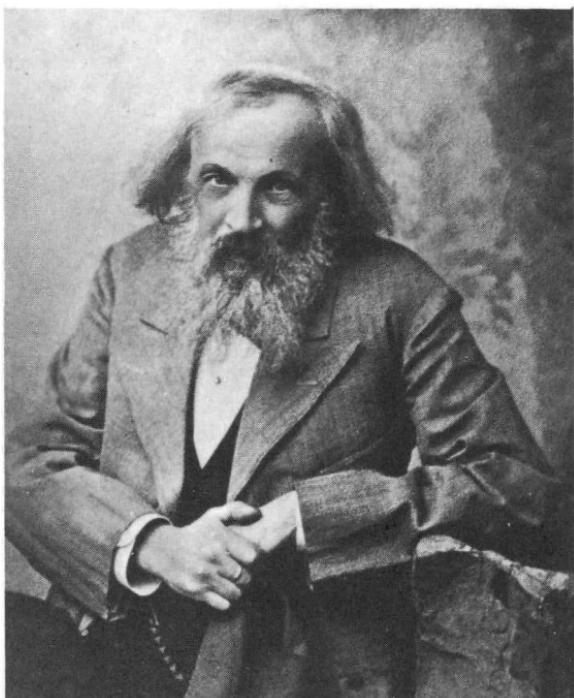
We can now readily understand why most alpha particles do not suffer appreciable scattering when they pass through the atoms in a thin sheet of metal. On the other hand, the occasional particles that come near the atomic nuclei are strongly scattered.

The Nuclear Charge and Atomic Number

Before 1913, all that was known about the number of positive charges on an atomic nucleus was that it is approximately half the atomic weight of the atom on the

conventional mass scale. In principle, the nuclear charge could have been determined more accurately, as it actually was some years later, from alpha-particle scattering. But the apparatus used by Geiger and Marsden was not capable of giving results of the required accuracy. Since good experimental values were not available, scientists turned to speculation; and in 1913, the Dutch physicist Antonius van den Broek put forward an idea which turned out to be correct.

During the decade 1860–1870, several scientists, notably Dmitri I. Mendeleev in Russia, showed that when elements are arranged in order of increasing atomic weights, with some minor exceptions, a definite periodicity (or repetition) in properties is apparent. This discovery led to the development of the *periodic table* in which elements are arranged in order of increasing weights; elements with similar chemical properties appear in the same vertical column (see the figure on the next two pages). Van den Broek suggested that the



Dmitri Mendeleev

PERIODIC CHART OF THE ELEMENTS

1 H															2 He		
3 Li	4 Be																
11 Na	12 Mg																
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57-71 La* Series	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103 Act Series	(104)	(105)													

“Lanthanide Series	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†Actinide Series	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

The periodic table in which the elements are arranged in order of increasing atomic weights (with a few exceptions). Elements with similar properties occur in the same vertical column

ALPHABETICAL LIST OF ELEMENTS AND SYMBOLS

Element	Symbol	Element	Symbol	Element	Symbol	Element	Symbol
Actinium	Ac	Erbium	Er	Mercury	Hg	Samarium	Sm
Aluminum	Al	Europium	Eu	Molybdenum	Mo	Scandium	Sc
Americium	Am	Fermium	Fm	Neodymium	Nd	Selenium	Se
Antimony	Sb	Fluorine	F	Neon	Ne	Silicon	Si
Argon	Ar	Francium	Fr	Neptunium	Np	Silver	Ag
Arsenic	As	Gadolinium	Gd	Nickel	Ni	Sodium	Na
Astatine	At	Gallium	Ga	Niobium	Nb	Strontium	Sr
Barium	Ba	Germanium	Ge	Nitrogen	N	Sulfur	S
Berkelium	Bk	Gold	Au	Nobelium	No	Tantalum	Ta
Beryllium	Be	Hafnium	Hf	Osmium	Os	Technetium	Tc
Bismuth	Bi	Helium	He	Oxygen	O	Tellurium	Te
Boron	B	Holmium	Ho	Palladium	Pd	Terbium	Tb
Bromine	Br	Hydrogen	H	Phosphorus	P	Thallium	Tl
Cadmium	Cd	Indium	In	Platinum	Pt	Thorium	Th
Calcium	Ca	Iodine	I	Plutonium	Pu	Thulium	Tm
Californium	Cf	Iridium	Ir	Polonium	Po	Tin	Sn
Carbon	C	Iron	Fe	Potassium	K	Titanium	Ti
Cerium	Ce	Krypton	Kr	Praseodymium	Pr	Tungsten	W
Cesium	Cs	Lanthanum	La	Promethium	Pm	(Wolfram)	
Chlorine	Cl	Lawrencium	Lr	Protactinium	Pa	Uranium	U
Chromium	Cr	Lead	Pb	Radium	Ra	Vanadium	V
Cobalt	Co	Lithium	Li	Radon	Rn	Xenon	Xe
Copper	Cu	Lutetium	Lu	Rhenium	Re	Ytterbium	Yb
Curium	Cm	Magnesium	Mg	Rhodium	Rh	Yttrium	Y
Dysprosium	Dy	Manganese	Mn	Rubidium	Rb	Zinc	Zn
Einsteinium	Es	Mendelevium	Md	Ruthenium	Ru	Zirconium	Zr

number of positive charges on the nucleus of an atom is equal to the ordinal number* of the element in the periodic system; this number was consequently called the *atomic number*. Thus, hydrogen (H), the first (lightest) element, has one positive charge on its nucleus; the next element, helium (He), has two; then comes lithium (Li) with three charges, beryllium (Be) with four positive charges, and so on up to 92 for uranium, as seen in the figure.[†]

Confirmation of the importance of the ordinal number of an element came later in 1913 from the work of Henry G.-J. Moseley performed in Rutherford's laboratory in Manchester, England. Some 2 years previously it had been found that each element can produce X rays which are characteristic of that element. Moseley was studying these characteristic X rays with a technique whereby lines were produced on a photographic plate; the position of each line was directly related to the wavelength (or energy) of a particular X ray. The results obtained for a series of consecutive elements, from calcium through zinc—with the exception of scandium—are shown in the historic photograph reproduced on page 32.[‡] It is apparent that the wavelengths of the characteristic X rays change in a regular manner with increasing atomic (ordinal) numbers of the elements.

From calculations made on the positions of the lines of the characteristic X rays, Moseley concluded that "there is in the atom a fundamental quantity which increases by regular steps as we pass from one element to the next (in the periodic table.) This quantity can only be the nuclear charge." Thus, from experimental data, Moseley identified the number of positive charges on the nucleus of an atom of

*The number of the element's position on the chart on page 28.

†The elements with ordinal (atomic) numbers 43, 61, and 93 through 105 have been made artificially; thus only 90 elements occur naturally on earth.

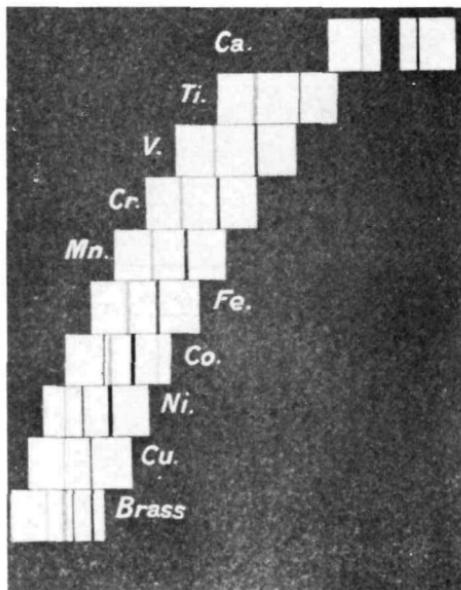
‡Brass (at the bottom of the figure) is an alloy of zinc and copper.



Henry G. J. Moseley

a given element with the ordinal (or atomic) number of the element in the periodic table.

During the years 1920–1925, several accurate measurements were made of alpha-particle scattering and the results confirmed the view that the nuclear charge is equal to the atomic (ordinal) number of the element. In fact, this concept is so well established that the atomic number of an element is now defined as the number of unit positive charges carried by a nucleus of an atom of that element. Originally the atomic number referred to the ordinal number of the



Moseley's historic photograph of the lines representing the characteristic X rays of consecutive elements from calcium to zinc (in brass). The lines of scandium, which follows calcium, are missing.

element in the periodic table; the latter number was then identified with the positive charge on the nucleus. This initial significance is still retained in the German language in which the word *Ordnungszahl*, meaning ordinal number, is used for the quantity we call the atomic number.

Composition of the Nucleus: Early Ideas

We have seen that the lightest element, hydrogen, has an atomic number of one; hence, the hydrogen nucleus carries a single positive charge. The hydrogen nucleus, now called a *proton*, * is, in fact, the unit of positive charge in the atom. It should be noted that the unit of positive electrical charge has exactly the same magnitude as the negative electronic charge. The unit negative charge, however, resides on a light particle,

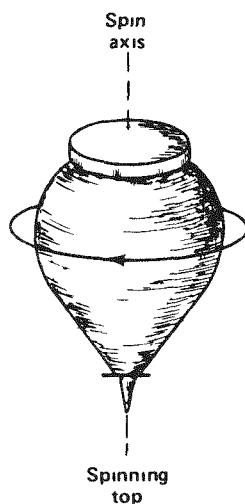
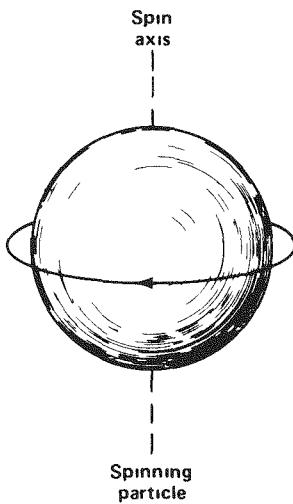
*From the Greek, *protos*, meaning "first".

the electron, whereas the unit positive charge is associated with a much heavier particle, the proton. As we have noted, the mass of a hydrogen atom, consisting of one proton and one electron, is 1837 times that of an electron. Hence, the proton is 1836 times as heavy as an electron. A positively charged electron (or *positron*) with the same mass as a negative electron is known, but it has a very transitory existence and is not a constituent of normal atoms.

The positive charges on an atomic nucleus are evidently due to the protons present. But is there any other kind of particle in the nucleus? We can easily show that there must be. The proton, being a hydrogen nucleus, has a mass of approximately one unit on the conventional atomic mass (or atomic weight) scale. Consequently, in order to account for the mass of a nucleus with an atomic weight of A , it would seem necessary to suppose that it contains A protons. But this cannot be so because the nuclear charge, and hence the number of protons, has been found to be about half (or less) of the atomic weight A .

At one time, scientists thought that the difficulty might be overcome by supposing that the nucleus of an element of atomic number Z and atomic weight A contained A protons (positively charged) plus $A - Z$ electrons (negatively charged). The net positive charge of the nucleus would then be the difference between the number of protons and the number of electrons, i.e., $A - (A - Z) = Z$; this is equal to the atomic number, as required. Since the $A - Z$ electrons would contribute only a small amount to the total mass, the mass of the nucleus would be determined essentially by the number of protons and would consequently be roughly equal to A units.

The view that electrons, as well as protons, were present in atomic nuclei appeared to receive support from the fact that the nuclei of many radioactive elements emit electrons; electrons produced in this manner are called *beta particles*. Nevertheless, there were difficulties associated with the idea.



Electrons protons neutrons and some nuclei have a motion called spin which is equivalent to that of a spinning top

One of the most outstanding was in connection with the "spin" of the nitrogen nucleus

Many nuclei (and other particles) appear to have a kind of motion resembling that of a top rotating (or spinning) about its axis (see the figure above). This spin is expressed in terms of a unit which we will describe later (p 58). Several different lines of experiment have established that the nucleus of nitrogen with an atomic weight of 14 ($A = 14$) has a spin of 1 unit. The atomic number of nitrogen is 7, i.e., $Z = 7$ (see the figure on page 28), and so if the theory of nuclear composition given above is correct, the nucleus should contain $A = 14$ protons and $A - Z = 7$ electrons. Protons and electrons each have either $+\frac{1}{2}$ or $-\frac{1}{2}$ unit of spin.* Consequently, the resultant spin of the total number of 21 protons and electrons in the nucleus could not possibly

*The + and - signs represent spins in opposite directions (see the figure on page 60)

be 1 unit. There is no way in which an odd number of particles with spins of $+\frac{1}{2}$ or $-\frac{1}{2}$ could be combined to give a net spin of 1 unit.

Neutrons in the Nucleus

There was no simple solution to this problem until 1932 when James Chadwick in England identified a new particle, called the *neutron*. The mass of the neutron, like that of the proton, is approximately 1 mass unit on the usual atomic weight (or mass) scale but, as its name implies, it is electrically neutral and carries no charge. Immediately after the discovery of the neutron was announced, Werner Heisenberg in Germany suggested that the nucleus contains protons and neutrons, but no electrons. This view is now universally accepted, because it accounts in a very satisfactory manner for the properties of atomic nuclei.

Since both protons and neutrons have a mass of about 1 unit, it follows that the nucleus of an atom of atomic weight A must contain a total of A protons and neutrons. Of these, only the protons provide the positive electrical charge, and so there must be Z protons in the nucleus of atomic number Z . The number of neutrons is consequently $A - Z$. Hence, the composition of any nucleus is as follows:

$$\text{Number of protons} = Z$$

$$\text{Number of neutrons} = A - Z$$

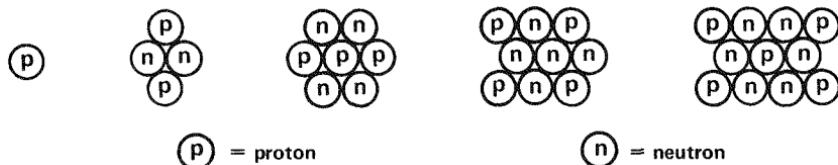
$$\text{Number of protons} + \text{neutrons} = A$$

Because they are constituents of the nucleus (and for other reasons), protons and neutrons are often referred to as *nucleons*. The total number of nucleons, A , in any nucleus is called the *mass number* of the atom. It is the whole number (or integer) closest to the atomic weight of a particular atomic species; the atomic mass or atomic weight itself is usually not a whole number.

The compositions of a few of the simpler atomic nuclei are shown in the figure. In ordinary hydrogen (H), which has



James Chadwick



Hydrogen	Helium	Lithium	Beryllium	Boron
$A = 1$	$A = 4$	$A = 7$	$A = 9$	$A = 11$
$Z = 1$	$Z = 2$	$Z = 3$	$Z = 4$	$Z = 5$

Compositions of some simple nuclei. (The shapes shown have no significance.)

mass number $A = 1$ and atomic number $Z = 1$, the nucleus is a single proton. In helium (He), $A = 4$ and $Z = 2$; hence there are two protons and two neutrons in the nucleus. And so on for the other nuclei represented in the figure. It should be noted that these particular arrangements of the protons and neutrons and the various resulting shapes are used simply for convenience. Only the numbers of protons and neutrons have any significance.

We may now reconsider the problem of the spin of the nitrogen nucleus in the light of the current theory of nuclear composition. The neutron, like the proton and electron, can have a spin of $+\frac{1}{2}$ or $-\frac{1}{2}$ unit. Since A is 14 for the nitrogen nucleus, it must contain 14 particles, 7 protons and 7 neutrons, each with a spin of $+\frac{1}{2}$ or $-\frac{1}{2}$. Because the total number of particles is even, it is possible for the individual spins to combine in such a way as to lead to a resultant spin of 1 unit. Thus, the spin of the nitrogen nucleus now presents no problem.

On the other hand, if there are no electrons in the nucleus, how is it possible to account for the emission of electrons (as beta particles) from radioactive nuclei? It turns out that a satisfactory explanation of the complex phenomena of beta-particle emission is to suppose that a neutron in the nucleus turns into a proton plus an electron, the latter being expelled as a beta particle. In this process

there is a conservation of mass and electrical charge, as indicated below.

Neutron \rightarrow Proton + Electron (beta particle)			
Mass	About 1	About 1	Very small
Electrical charge	0	+1	-1

However, the spins do not balance in the conversion of a neutron into just a proton and an electron. It is impossible for the $+\frac{1}{2}$ or $-\frac{1}{2}$ spin of a neutron, on the left side, to balance the spins of the proton and electron, each of which is $+\frac{1}{2}$ or $-\frac{1}{2}$, on the right side. This discrepancy is accounted for by the simultaneous emission of another particle called a *neutrino*; it has no mass (when at rest) and no charge, but it does have $+\frac{1}{2}$ or $-\frac{1}{2}$ unit of spin. The existence of the neutrino, which also explains other beta-particle phenomena, has been confirmed by experiment.*

Isotopes

Let us return once more to the general problem of the structure of the atom. Since the nucleus contains Z protons, that is, it has Z unit positive charges, there must be Z electrons, with Z unit negative charges, in the space around the nucleus. The atom as a whole is then electrically neutral. It is the Z orbital electrons that determine the chemical behavior and many of the physical properties, such as magnetism and optical and X-ray spectra, of the element. The characteristic feature that makes one element different from another is thus the atomic number, which is the number of protons in the nucleus of the atom of the element, or the number of electrons in the space around the nucleus. Both numbers must be equal.

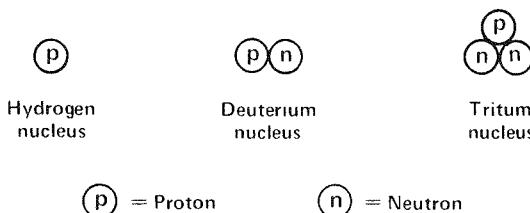
*For more information on the neutrino, see *The Elusive Neutrino*, another booklet in this series.

One implication of the foregoing conclusion is that nuclei of the same element may have different mass numbers. The atoms of the given element will then have different atomic weights (or atomic masses); but, provided the nuclei contain the same number of protons, the atoms are those of the same element. If the numbers of neutrons are different, the masses will differ. Forms of the same element with atoms of different mass are called *isotopes* of that element. For example, as will be considered more fully below, there are three well known isotopes of the element hydrogen with masses (mass numbers) of 1, 2, and 3, respectively. Similarly, oxygen exists in the air and in water as three isotopes with masses 16, 17, and 18, respectively. It happens that in each case the lightest isotope is by far the most common, but this is not always so.

The term isotope is derived from two Greek words meaning "the same place" (in the periodic table of the elements). Isotopes were first identified from the study of radioactive elements in 1913, long before the present theory of nuclear composition had been developed. It was then found that two or more elementary species, having different atomic masses, would have to occupy the same position in the periodic table; from this came the word isotope. The reason for the existence of the isotopes was not well understood at the time, but it is now known that isotopes of an element have the same number of protons but different numbers of neutrons in their atomic nuclei.

Consider, for example, the element hydrogen; as mentioned above, it exists in three isotopic forms. The atomic nuclei of each isotope contain one proton, so that the atomic number Z is 1 in each case. The lightest isotope, commonly referred to as hydrogen, symbol H, has $A = 1$. Thus there are no neutrons in its nucleus and it is, in fact, the only atomic nucleus that does not contain any neutrons. The next isotope is called deuterium, symbol D, from the Greek word meaning the "second"; its mass number A is 2, so it has one neutron

in its nucleus. Finally, there is tritium, symbol T, from the Greek for the “third”; for this isotope $A = 3$ and the nucleus contains two neutrons. The compositions of the atomic nuclei of the three isotopes of hydrogen are illustrated in the figure.



Compositions of the nuclei of the three isotopes of hydrogen; each contains 1 proton, but the numbers of neutrons are different. .

In 1929 scientists discovered that the familiar element oxygen ($Z = 8$) occurs in nature in three isotopic forms with mass numbers $A = 16, 17$, and 18 , respectively. The isotopes are called oxygen-16, oxygen-17, and oxygen-18 and are represented by the symbols ^{16}O , ^{17}O , and ^{18}O . The nuclei thus contain 8 protons and either 8, 9, or 10 neutrons.

For many years, chemists had expressed the atomic weights of the elements on a scale in which the atomic weight of oxygen in nature had been set at exactly 16.0000. This meant that the average of the masses of the three isotopes, weighted in accordance with their known proportions, was taken to be 16.0000. On the other hand, in physical determinations of atomic masses by means of the mass spectrometer, the single isotope oxygen-16 was assigned the mass 16.0000. The weighted average for the three isotopes on this basis turned out to be 16.0045. Consequently, some confusion resulted from a situation in which atomic weights (or masses) had different values depending on whether they were stated on the chemical or the physical mass scale.

The Atomic Mass Unit

The situation was resolved on January 1, 1962, by an international agreement among physicists and chemists to adopt the same reference value for expressing atomic masses. The atomic weight (or mass) of the carbon-12 isotope, the main isotope of the element carbon, is taken to be exactly 12.0000 units. Atomic weights of the elements as they occur in nature as mixtures of isotopes and the atomic weights of individual isotopes, commonly referred to as *atomic masses*, are now expressed on the carbon-12 scale.

The unit of atomic mass (or the *atomic mass unit*), abbreviated to 1 amu, is thus one-twelfth of the mass of the atom of the carbon-12 isotope. Its value in more familiar units is 1.67×10^{-24} gram. The precise masses of the proton, neutron, and electron are then as follows:

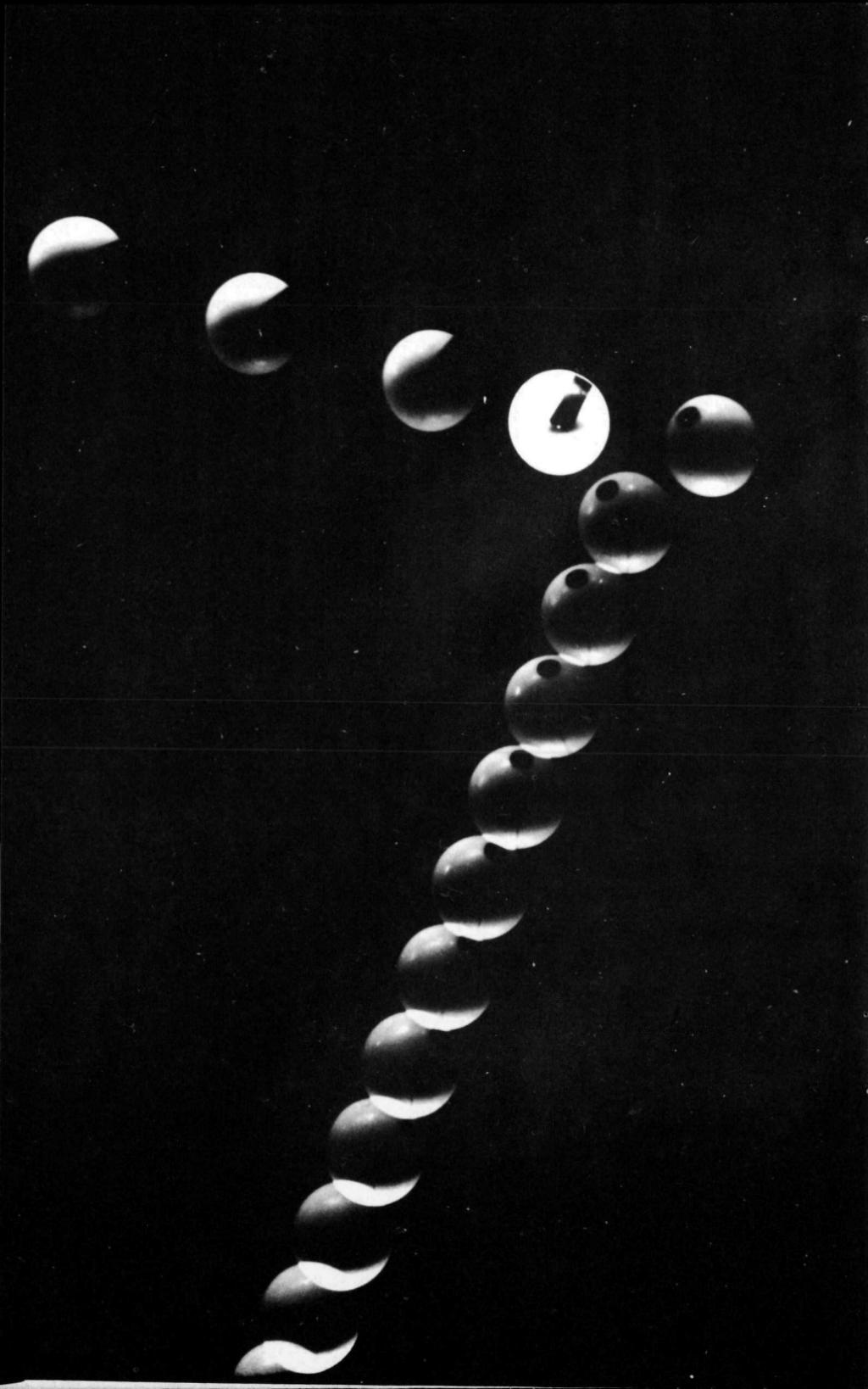
$$\text{Mass of proton} = 1.007277 \text{ amu}$$

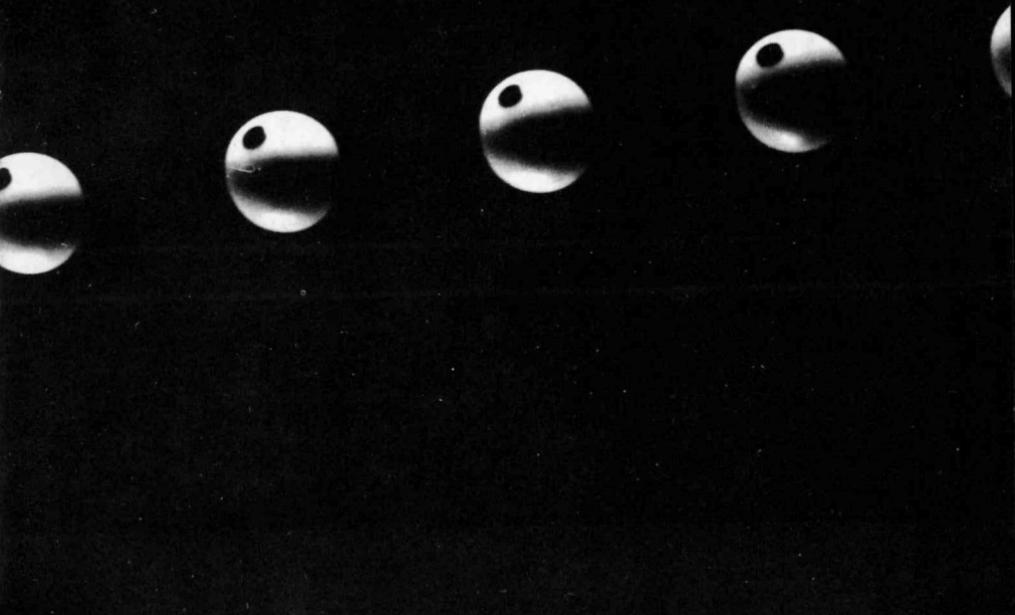
$$\text{Mass of neutron} = 1.008665 \text{ amu}$$

$$\text{Mass of electron} = 0.000549 \text{ amu}$$

Thus the masses of both proton and neutron are somewhat greater than 1 amu. We might expect, therefore, that since nuclei contain protons and neutrons, their masses would be significantly larger than a whole number. But this is not so. In fact, the masses of nuclei of mass number equal to or greater than 10 are less than a whole number and for all nuclei the masses are less than the sum of the masses of the constituent protons and neutrons. The significance of this apparently surprising fact will be considered later (p. 78).

Of the 90 elements known to exist on the earth, all but 20 occur in nature as mixtures of two or more isotopes. Tin, for example, has as many as ten isotopes and mercury has seven. Unstable (radioactive) isotopes of the 90 natural elements and of 15 artificially produced elements have been created by various nuclear reactions. At present, some 1500 isotopes have been identified; of these, about 280 are found on the earth.





It is important to realize that a physical theory—even a very profound one like relativity or the quantum theory—often has a very simple, almost pictorial idea at its center. The great physicists have all been masters of what has come to be known as “Gedanken experiments”—experiments existing only in the mind, unperformed and probably unperformable, which when contemplated, put physical ideas into new juxtapositions. For example . . . Heisenberg, in the late 1920s, imagined a microscope through which he could try to measure the position of an electron in an atom, and then realized that the light quanta reflected from the electron into the microscope would inevitably change the momentum of the electron—a realization that is at the heart of the Heisenberg uncertainty principle.

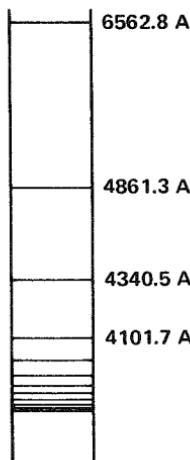
Jeremy Bernstein

THE ORBITAL ELECTRONS

Atomic Spectra

We will say more about the nucleus later, but now some consideration will be given to the orbital electrons in the atom. The number of these electrons is known for the atoms of any element, since it is equal to the atomic number of the element. But how are these electrons arranged in the space around the nucleus? Much information on this subject has been obtained from a study of atomic spectra.

When a gas (or a vaporized liquid or solid) is heated sufficiently or is stimulated by an electrical discharge, it often emits radiations (or rays) in the ultraviolet, visible, or infrared range. If the rays are examined with a *spectroscope* (which is an instrument for splitting up complex radiation into its components) a definite pattern of lines, known as a *spectrum*, appears.* An example of such a line spectrum is given in the figure.



Part of the line spectrum of the hydrogen atom. The wavelengths of four lines are given in angstrom units ($\text{\AA} = 10^{-8} \text{ cm}$).

*See *Spectroscopy*, a companion booklet in this series for a description of radiations and spectra.

Each element produces its own characteristic line spectrum, usually referred to as an *atomic spectrum* because it originates in the atoms—actually from the electrons—of the element. The lighter atoms, such as hydrogen and helium, with a small number of electrons, yield fairly simple spectra with relatively few lines, which can be readily identified. But for some of the heavier atoms, the spectra may consist of hundreds of lines and their analysis is difficult.

Wavelength and Frequency of Radiation

Light can be regarded as a wave motion, and each line in a spectrum represents a particular *wavelength* (the distance from one wave crest to the next one). (See the figure on the next page.) The wavelength is related to the *frequency*, or number of waves passing a given point per unit of time. The wavelength, indicated by the symbol λ , and the frequency, indicated by ν , are related in the following manner. If c is the velocity of light, expressed in centimeters per second, a wave will travel c cm in 1 sec. Hence, c/λ waves will pass by a given point in 1 sec, where λ is the wavelength in centimeters. By definition, this is equal to the frequency, with time in seconds; thus,

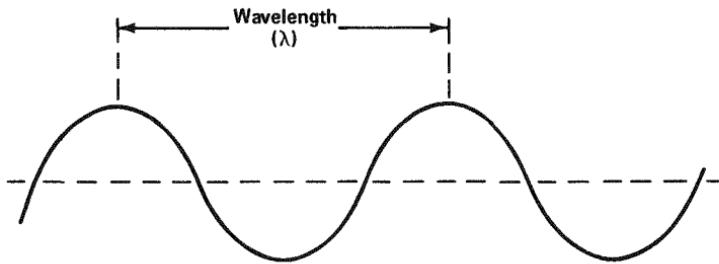
$$\nu = \frac{c}{\lambda} \text{ or } \lambda = \frac{c}{\nu} \quad (1)$$

where ν = frequency in waves (or cycles) per second

c = velocity of light in centimeters per second

λ = wavelength in centimeters

The velocity of light (in a vacuum) is constant, roughly 3×10^{10} cm/sec, regardless of the wavelength. Hence, the frequency for a given spectral line is inversely proportional to the wavelength of the radiation producing that line; that is to say, the frequency is large when the wavelength is small and vice versa.



The wavelength is the distance between two successive crests of a wave.

Quantum Theory

Another way in which a spectral line can be characterized is in terms of energy. The relationship between the frequency of a given radiation and a quantity of energy is based on the *quantum theory* of radiation. According to this theory, which was first proposed by Max Planck in Germany in 1900 and extended by Albert Einstein in Switzerland in 1905, a body cannot take up or give out energy in the form of radiation in a continuous manner. The energy can be absorbed or emitted only as integral (whole number) multiples of a definite amount called a *quantum*. The energy of the quantum is related to the frequency of the radiation by

$$E = h\nu \quad (2)$$

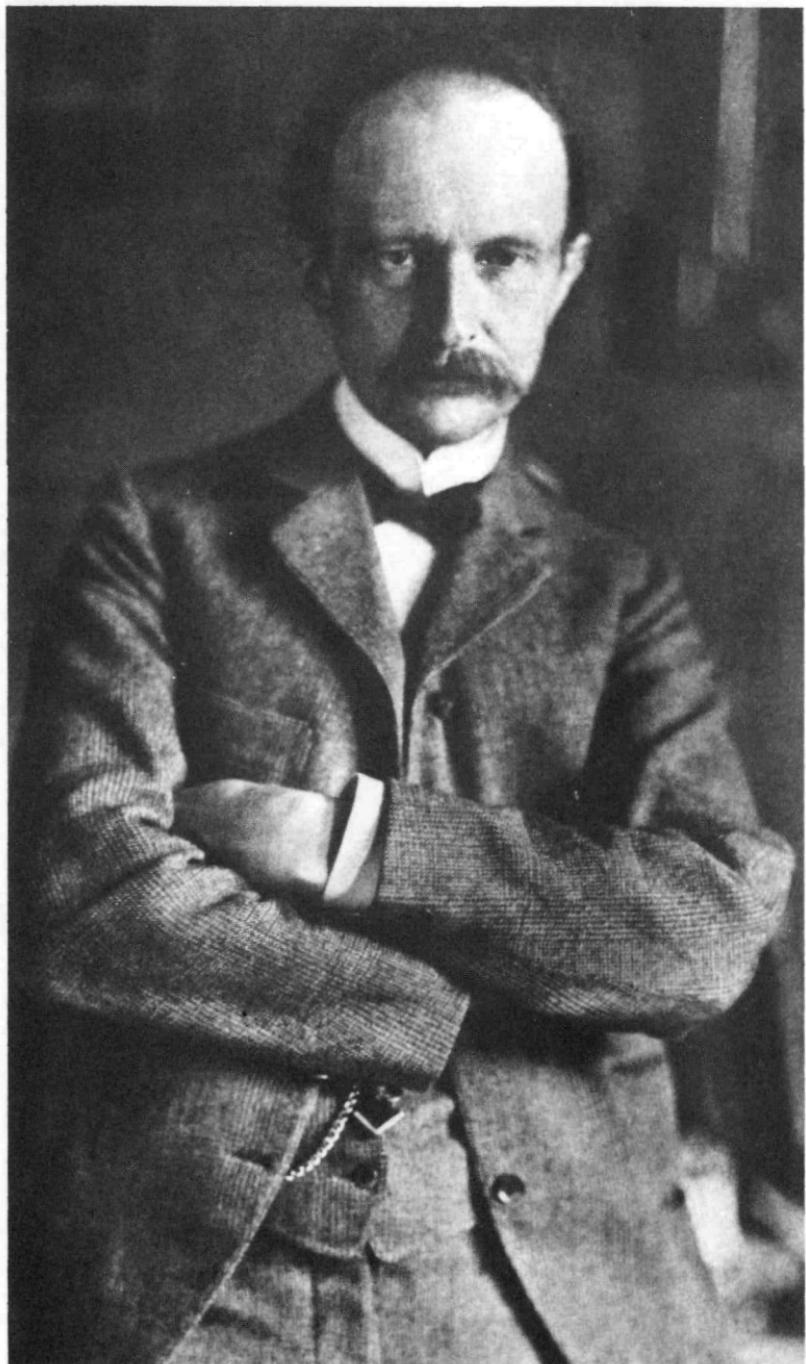
where ν = frequency

E = energy

h = a universal constant (Planck's constant)

The Origin of Spectra

In order to explain why the negatively charged orbital electrons did not fall into the positively charged nucleus,



Max Planck

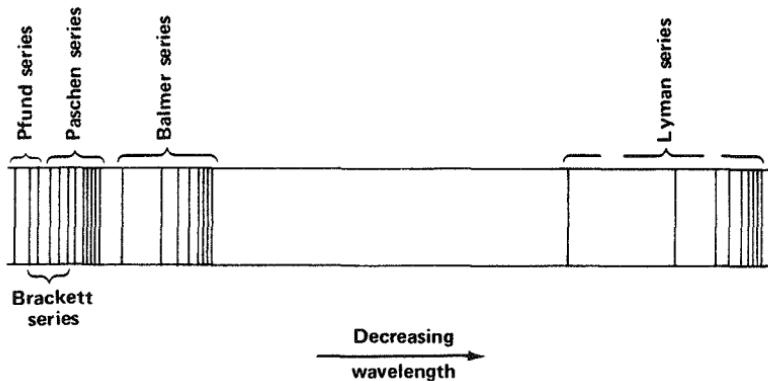
Rutherford postulated that the electrons revolved rapidly in orbits about the nucleus, just as the planets revolve about the sun. The inward attractive force of the positive nucleus would then be balanced by the outward centrifugal force on the revolving electrons. However, this analogy between an atom and the sun's planetary system is false, because of the presence in atoms of electrically charged particles.

By the electromagnetic theory of light (and similar radiations), propounded by the Scottish mathematical physicist James Clerk Maxwell in 1864, an electron revolving in an orbit should emit energy continuously as radiation. But if this were so, and Newton's laws of motion applied as they do for planetary motion, the radius of the electron orbit should decrease steadily because of its loss of energy. This would result in a continuous change in the character of the radiation emitted. The spectrum of an atom would then cover a continuous range of wavelengths instead of consisting, as it does, of well-defined lines.

A historic step toward resolving this difficulty and understanding atomic spectra, as well as the arrangement of the electrons in an atom, was taken in 1913 by the Danish physicist Niels Bohr in Rutherford's laboratory in Manchester, England.

The Spectrum of Hydrogen

Experimental observations of atomic spectra in general have revealed certain regularities. In particular, it has been found that lines in the spectrum fall into a number of groups called *series*. Five such series are known in the spectrum of atomic hydrogen; they are indicated, along with the names of their discoverers, in the figure. The Lyman series is in the ultraviolet region (short wavelength) and the Balmer series is in the visible range; the other three series are in the infrared (long wavelength) region. Although in 1913 these five series in the hydrogen spectrum had not all been identified, it had been established from actual measurements that the



Lines in the five known series in the spectrum of atomic hydrogen. In the Pfund, Brackett, and Paschen series some of the lines overlap.

frequency of any known line could be represented by the general formula

$$\nu = \frac{Rc}{x^2} - \frac{Rc}{y^2} \quad (3)$$

where R = a universal constant (Rydberg constant)

x = an integer (1, 2, 3, etc.)

y = an integer (2, 3, 4, etc., but always larger than x)

ν = the frequency

c = velocity of light

The values of x and y for the five series in the spectrum of hydrogen are given in the following table.

Series	x	y
Lyman	1	2, 3, 4, ...
Balmer	2	3, 4, 5, ...
Ritz-Paschen	3	4, 5, 6, ...
Brackett	4	5, 6, ...
Pfund	5	6, 7, ...

By the relationship given above, the observed frequencies of the lines in any one series of the spectrum of the hydrogen



Niels Bohr

atom can be expressed as the difference of two quantities; one of these quantities, Rc/x^2 , is constant in each series whereas the other, Rc/y^2 , decreases regularly from one line to the next as y increases one unit at a time. This is shown by a regularly decreasing set of line spacings in terms of wavelength (see the figure on page 49)

Bohr's Theory of Atomic Spectra

To account in a theoretical manner for the observed spectrum of hydrogen (and of other atoms), Bohr applied quantum theory to Rutherford's idea of an atom consisting of a central nucleus with electrons revolving around it in circular orbits. But he postulated that the number of such orbits is restricted by the requirement that the angular momentum* of the electron be "quantized". That is to say, the angular momentum of the electron must be a whole number of multiples of a certain quantity, $h/2\pi$, where h is Planck's constant. Stated in another way, the first of Bohr's postulates is that the angular momentum Mvr of an electron of mass M moving with velocity v in a circular orbit of radius r is always given by $nh/2\pi$, where the quantum number n is an integer, i.e., 1, 2, 3, 4, etc. These permissible orbits were called *stationary orbits*. The requirement for such orbits, by Bohr's first postulate, was thus

$$Mvr = n \frac{h}{2\pi} \quad (4)$$

The second postulate is that, contrary to the requirement of the original electromagnetic theory of radiation, an electron does not emit radiation while moving in a stationary orbit. Thus, as long as the electron remains in a stationary orbit, its energy is unchanged; the value of the energy is then determined by the quantum number n described in the first

*The angular momentum of a particle moving in a circular orbit of radius r is Mvr , where M is the mass and v is the velocity of the particle.

postulate. For the hydrogen atom, for example, Bohr showed mathematically that the energy E of an electron in a given stationary orbit characterized by n can be expressed as

$$E = \text{Constant} - \frac{Rhc}{n^2} \quad (5)$$

Finally, Bohr said that the emission (or absorption) of radiation occurs only when an electron transfers from one stationary orbit to another. The difference in energy between the two orbits is then emitted (or absorbed) in definite quanta, as required by the quantum theory of radiation. Suppose an electron transfers from an initial stationary orbit in which its quantum number is n_i to a final stationary orbit with quantum number n_f ; this is often referred to as a “quantum jump” (see the figure). The energy E , in the initial state is given by equation (5) as

$$E_i = \text{Constant} - \frac{Rhc}{n_i^2}$$

whereas in the final state the energy E_f is

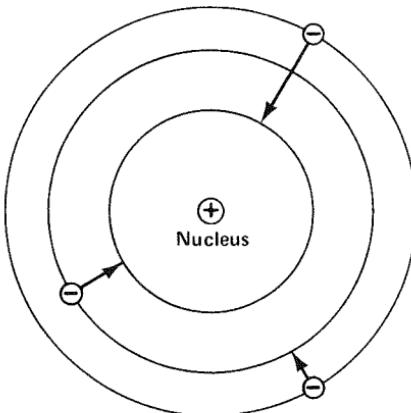
$$E_f = \text{Constant} - \frac{Rhc}{n_f^2}$$

The energy E emitted as a result of the quantum jump of an electron from any one stationary orbit to another is then

$$E = E_i - E_f = \frac{Rhc}{n_f^2} - \frac{Rhc}{n_i^2}$$

If this energy is emitted as radiation, the frequency is given by the quantum theory equation (2), $\nu = E/b$, as

$$\nu = \frac{E}{b} = \frac{Rc}{n_f^2} - \frac{Rc}{n_i^2} \quad (6)$$



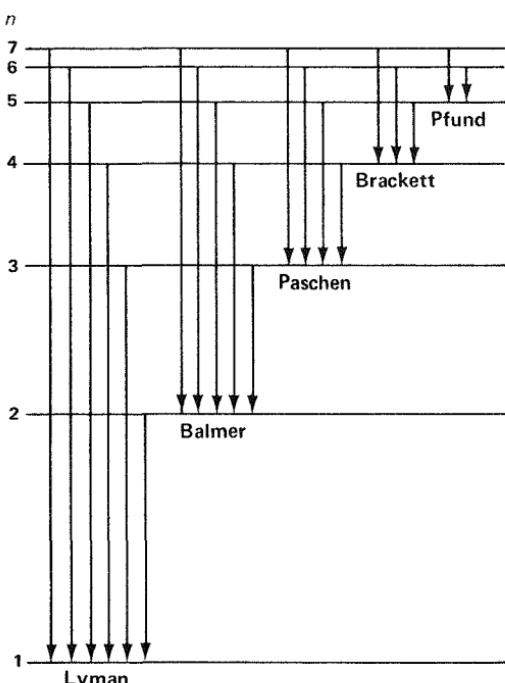
Possible quantum jumps (transitions) of electrons in Bohr orbits.

The production of a line of definite frequency (or wavelength) in the spectrum of hydrogen, or, indeed of any atom, can thus be attributed to the radiation of a definite amount of energy associated with the transition or "jump" of an electron from one stationary orbit (or state) to another stationary state.

In physical terms, we may think of a line spectrum as originating in the following manner. Consider an atom that is heated or is stimulated by an electrical discharge; it takes up energy and an electron is transferred to an outer stationary orbit in which its energy is higher than normal. After a very short time, the electron jumps into an inner stationary orbit of lower energy. The energy difference between the two states is emitted as a quantum of radiation, representing a particular line in the spectrum; the frequency of the line is then given by equation (6). Since there are several possible stationary orbits, the complete spectrum may contain many individual lines.

Comparison of the theoretical equation (6) with equation (3) for the actual (observed) frequencies of the lines in the hydrogen spectrum shows that the two equations are

fundamentally the same. Thus, Bohr was able to account for the general character of the spectrum of the hydrogen atom. Furthermore, the value of the Rydberg constant R calculated from the theory was in good agreement with that derived from experimental measurement of the frequencies (or wavelengths) of the lines in the hydrogen spectrum. The integers x and y in equation (3) are seen to represent the quantum numbers n_f and n_i of the final and initial orbits, respectively, of the electron. Since x is the same for all the lines in a given series, it follows that these lines arise from transitions in which the electrons start from orbits with different quantum numbers ($n_i = y$), but all end up in the same final orbit ($n_f = x$)



Electronic transitions associated with the five series in the spectrum of atomic hydrogen

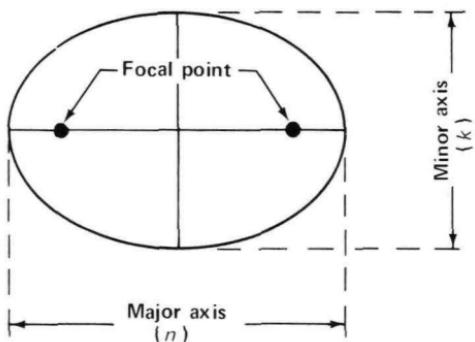
We shall see later that the concept of electrons moving in definite orbits is no longer acceptable. Consequently, the so-called stationary orbits, in which the electron has specific energy values, are now usually referred to as the *energy levels* or *energy states* of the electron. According to Bohr's theory, these levels (or states) are characterized by the value of the quantum number n . The positions of the first seven energy levels of the hydrogen atom ($n = 1, 2, 3, 4, 5, 6, 7$) are indicated in the figure. The arrows show the electronic transitions (or quantum jumps) observed in the five known series in the spectrum of hydrogen. For the Lyman series n_f (or x) is 1, for the Balmer series it is 2, and so on.

Bohr's theory was developed in detail to account for the spectrum of hydrogen but the general concept of definite electronic energy levels (or states) is applicable to all atoms. Each line in the spectrum is then attributed to a transition from one such electronic state to another. The frequency (or wavelength) of the line is related by quantum theory to the difference in energy between the final and initial states.

Sommerfeld's Elliptical Orbits

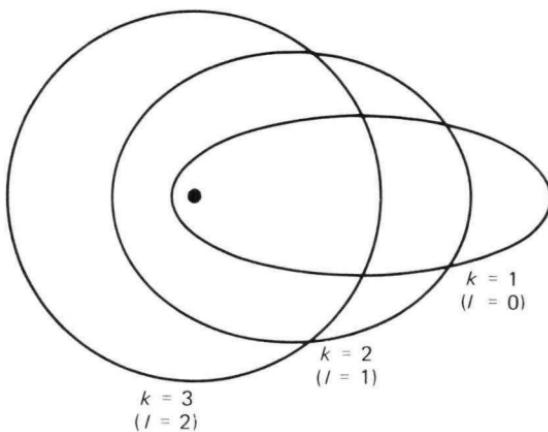
Although the idea of definite orbits is obsolete, and this should be borne in mind, it is nevertheless instructive to follow the historical development in which orbits are assumed to have reality. In the simple Bohr theory, the orbits were supposed to be circular; but in 1915, Arnold Sommerfeld in Germany showed that elliptical orbits should also be possible. The nucleus would then be located at one of the two focal points of the ellipse. To describe an elliptical orbit, two quantum numbers would be required: a principal quantum number n to define the energy and an orbital quantum number k which determines the orbital angular momentum. For any given value of n , the possible values of k would be $n, n - 1, n - 2, \dots, 1$; for example, if n is 4, then k could be 4,3,2,1. In physical terms, n is a measure of the major axis of the elliptical orbit and k represents the minor

axis (see the figure). When n and k are equal, the orbit is circular; the circle is a special (limiting) case of an ellipse in which the major and minor axes are equal and the two focal points converge to form the center of the circle.



Major and minor axes and focal points of an ellipse.

The shapes of the permissible electron orbits for $n = 3$ ($k = 3, 2, 1$), according to the Bohr-Sommerfeld theory, are shown in the figure; the heavy dot indicates the position of the atomic nucleus. The length of the major axis, which is equal to the diameter of the circular orbit ($n = 3, k = 3$), is the same for each orbit, but with decreasing k the orbit becomes more elliptical. In a circular orbit, the electron would always remain at the same distance from the nucleus. In an elliptical orbit the distance would change continuously; sometimes the electron would be nearer to the nucleus than in a circular orbit and at other times it would be farther away. The electron would be expected to travel more slowly when it is farther from the nucleus. Consequently, in an elliptical orbit the electron would spend more time farther from, rather than closer to, the nucleus. The significance of this point will be seen later in connection with the modern theory of the behavior of the orbital electrons.



The three permissible Bohr-Sommerfeld stationary orbits for the $n = 3$ electron state in the hydrogen atom; as k (or l) decreases the orbit becomes more elliptical.

Sommerfeld developed the theory of elliptical orbits to account for the occurrence in some hydrogen (and related) spectra of a group of closely spaced lines where a single line was to be expected from Bohr's theory. For the present purpose, however, the main point is that, in order to explain the details of spectral lines (their fine structure), it was necessary to introduce a second quantum number for the electrons in an atom. In later developments, such a quantum number is still required, but its possible values are $n - 1$, $n - 2$, $n - 3$, . . . , 0. That is to say, for $n = 4$, for example, the second quantum number can be 3, 2, 1, 0, rather than 4, 3, 2, 1 as required by Sommerfeld's theory. The second quantum number is consequently represented by l in order to avoid confusion with Sommerfeld's k ; actually, l is equivalent to $k - 1$. From now on, however, the symbol l will be used.

Magnetic and Spin Quantum Numbers

An electrically charged particle moving in an orbit behaves like a magnet. Hence, an electron in an atom is

equivalent to a magnet; and the direction of its orbit would be affected by a magnetic field. In fact, the Dutch physicist Pieter Zeeman had observed in 1896 that the spectral lines of an atom are often split into closely spaced components when the emission of radiation takes place in a magnetic field. This result indicated the necessity of adding a third quantum number m , called the *magnetic quantum number*, to account for the behavior of the electrons in an atom.

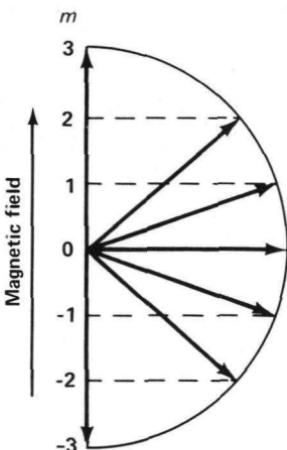
To allow for the magnetic properties of an electron in an orbit, it is assumed that the geometric projection of the orbital angular momentum in the direction of the magnetic field is quantized. In other words, the permitted directions of the angular momenta are such that the projections of the lines representing these momenta are whole-number multiples of $h/2\pi$. This whole number is the magnetic quantum number m . For any given value of l , there are $2l + 1$ possible values of m , namely $l, l - 1, l - 2, \dots, 0, \dots, -(l - 2), -(l - 1), -l$.

For example, if $l = 3$, the $2l + 1 = 7$ permissible values of m are $3, 2, 1, 0, -1, -2, -3$; this situation is illustrated in the figure on page 60. The heavy arrows show the seven permitted directions of the orbital angular momentum of an electron in a magnetic field. The corresponding values of the magnetic quantum number m are given at the left by the projections along the direction of the magnetic field. It will be noted that the values of m can be negative as well as positive.

We saw earlier that an electron appears to behave as if it were rotating (or spinning) about its own axis. This idea was introduced in 1925 by George E. Uhlenbeck and Samuel Goudsmit in Holland; they postulated the existence of a spin quantum number s in order to explain the splitting of some spectral lines into two distinct components. The spin angular momentum of an electron is always $\frac{1}{2}$ in terms of the usual $h/2\pi$ quantum of angular momentum. Since an electron can spin about its axis either in one direction or in the other (see the figure on page 60), the possible values of s are $+\frac{1}{2}$ or $-\frac{1}{2}$. These are the only permissible values of the spin



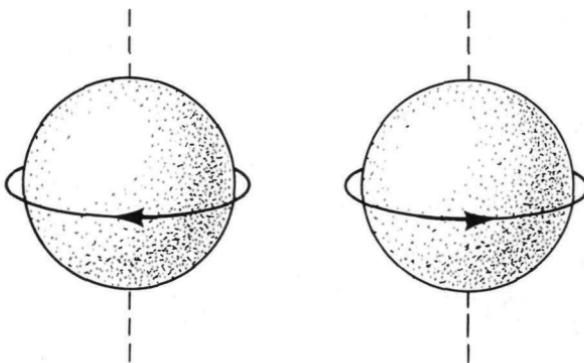
Pieter Zeeman



The seven permissible values of the magnetic quantum number (m) of an electron when the orbital quantum number is $l = 3$.

quantum number, regardless of the values of the other three quantum numbers, n , l , and m .

With the introduction of the four quantum numbers of an electron we are in a position to consider an important aspect of the structure of atoms, namely, the arrangement of the electrons in the space around the nucleus. But before doing so, it is desirable to say something about the meaning of electron orbits.



An electron can spin about its axis either in one direction or in the other; there are only two possible values of the spin quantum number, $+1/2$ or $-1/2$.

Electrons as Waves

In order to explain a number of puzzling physical phenomena, Heisenberg (see p. 35) in 1927 introduced the *uncertainty principle*. For the present discussion, this principle may be stated in the following form: It is impossible to determine simultaneously the exact position and momentum of any particle; if the position of the particle is known exactly, then the momentum is completely uncertain, and vice versa. But in the theories of Bohr and of Sommerfeld the angular momentum of an electron is assumed to be exactly known when it is in a particular stationary orbit, i.e., at an exactly known position. Consequently, if the uncertainty principle is correct, and there are many good reasons for thinking that it is, then the postulate of stationary orbits must be wrong. Yet, with some small modifications, the general conclusions reached by Bohr and Sommerfeld are correct. Why is this so?

The first clue was provided in 1923 by Louis Victor de Broglie in France, who reasoned along the following lines. Let us consider light and similar radiations. For many years, light has been (and still is) regarded as a wave motion because many of its properties can be explained in this manner. On the other hand, some of the properties of radiation require that it be treated as a particle, called a *photon*, which carries the quantum ($h\nu$) of energy. Thus, radiation exhibits a *wave-particle duality*; some of the properties of radiation are characteristic of waves whereas others are typical of particles. The suggestion made by de Broglie was that the wave-particle dualism is applicable to material particles, as well as to radiation. This dualism is, in fact, a manifestation of the more general uncertainty principle referred to above.*

*For more about wave-particle duality and the uncertainty principle, see *Microstructure of Matter*, a companion booklet in this series.



Werner Heisenberg



Louis Victor de Broglie

In developing the idea of wave-particle duality, de Broglie turned to the famous mass-energy equation derived by Einstein in 1905. An aspect of the theory of relativity is that mass and energy are equivalent; in other words, mass may be regarded as a form of energy. The relationship between mass and energy is given by

$$E = Mc^2 \quad (7)$$

where E = energy (ergs)

M = mass (grams)

c = velocity of light (3×10^{10} cm per sec)

The Einstein equation (7) gives the energy equivalent E of a particle of mass M . If the particle is a photon of radiation, carrying a quantum of energy, this energy can be represented by the Planck equation (2), $E = h\nu$. If the two expressions for E are set equal to each other, we see that

$$Mc^2 = h\nu$$

or, written in another way,

$$Mc = \frac{h\nu}{c}$$

According to equation (1), c/ν is equal to λ , the wavelength of the radiation, so that

$$Mc = \frac{h}{\lambda} \quad (8)$$

The result just obtained follows from the wave-particle duality of radiation in which the particle (the photon) travels with the velocity of light c . If a similar relationship is to apply to material particles traveling with a velocity v , then,

argued de Broglie, c in equation (8) should be replaced by v , to give

$$Mv = \frac{h}{\lambda}$$

or

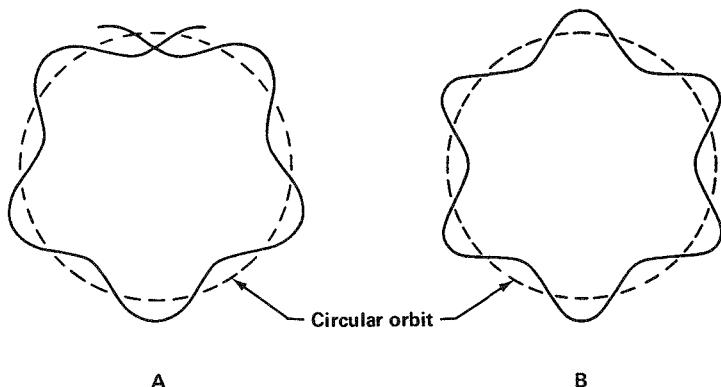
$$\lambda = \frac{h}{Mv} \quad (9)$$

This is the de Broglie equation for the length λ of the wave associated with a particle of mass M traveling with a velocity v . The association of wave properties with material particles, such as electrons, protons, and neutrons, has been established by experiment and the lengths of the “matter waves” are in agreement with the de Broglie equation.

It should be mentioned that the expressions “wave properties” and “matter waves” must not be taken to imply that there are actual waves. A possible physical interpretation will be considered later. But for the moment it is sufficient to say that subatomic particles must be treated mathematically by the same equations as apply to waves.

We can now examine the significance of Bohr’s stationary orbits. Instead of treating the electron as a particle moving in an orbit about the nucleus, de Broglie said that we should think of a series of waves traveling around an orbit. If the wavelength is such that the waves overlap, as they do in figure (A), the waves will eventually cancel out; such an orbit would have no physical significance. If an orbit is to be a stationary one, it must contain a whole number of wavelengths, as in figure (B) on the next page.

The total length of a circular orbit of radius r is $2\pi r$, and the number of waves in it are obtained upon dividing by the wavelength λ ; the result is $2\pi r/\lambda$. This must be equal to an integer n to give a permissible orbit, so that



In (A) the waves overlap and will eventually cancel out; the stationary circular orbit in (B) contains a whole number of wavelengths.

$$\frac{2\pi r}{\lambda} = n$$

is the required condition. The de Broglie value b/Mv from equation (9) may now be substituted for λ , and upon rearranging, the result is seen to be

$$Mvr = n \frac{b}{2\pi}$$

This is identical with equation (4) derived from Bohr's first postulate. The quantization of angular momentum in terms of units of $b/2\pi$, which Bohr had to *assume* in order to obtain correct results, is thus seen to be a natural consequence of the wave behavior of electrons.

Wave Mechanics

The next step was taken in 1926 by the Austrian theoretical physicist Erwin Schrödinger who applied mathematical procedures to the wave-particle duality concept of



Erwin Schrödinger

matter. This led to the development of *wave mechanics*, which has been outstandingly successful in interpreting many properties of the atom.* Although it is not the only possible interpretation, we may think of the mathematical treatment somewhat along the following lines.

The position of an electron (or other particle) of definite momentum or energy cannot be known exactly because of the uncertainty principle. All that is possible is to state the statistical probability that the particle will be found at any given point. Wave mechanics then postulates that this probability can be expressed by a mathematical equation similar to that used for describing wave behavior in general. In other words, the statistical probability of finding a particle in a specified position can be represented by an equation of the same form as that which describes the amplitudes of waves, for example, sound waves, in space.

By applying the methods of wave mechanics to the electrons in an atom, certain equations are obtained. These equations are found to have solutions only for *integral values* of quantities which may be represented by the symbols n , l , and m . Thus, the three quantum numbers introduced somewhat arbitrarily into the Bohr-Sommerfeld theory emerge as a necessary consequence of the wave equations. One difference, however, is that the angular momentum is not given by l (or s) times $h/2\pi$, but rather by $\sqrt{l(l+1)} h/2\pi$. The spin quantum number is not a requirement of ordinary wave mechanics, but it is found to be needed when Einstein's theory of relativity is introduced.

Orbitals Replace Orbits

Not only are the quantum numbers, n , l , and m , a direct requirement of wave mechanics, but they are an im-

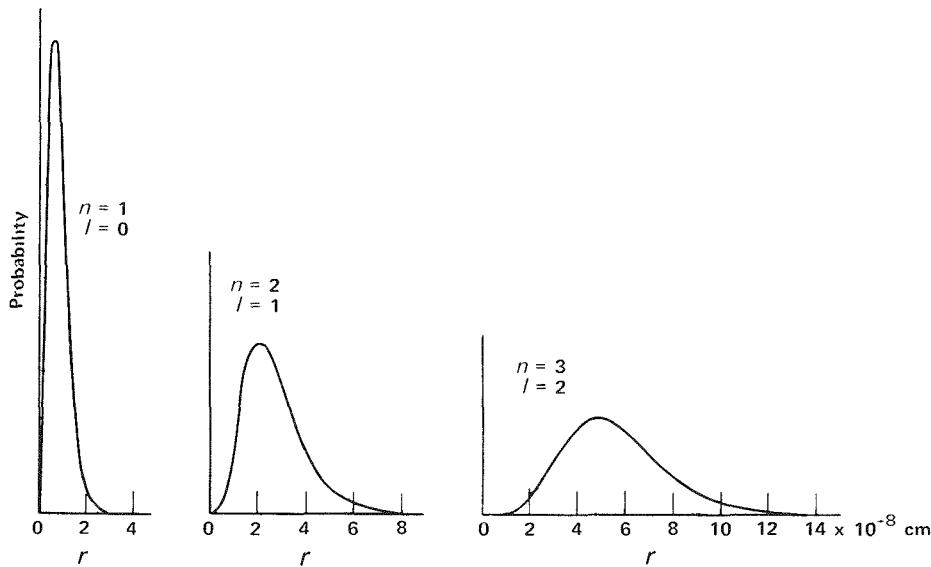
*Another mathematical approach, called *matrix mechanics*, was developed at about the same time. Both are aspects of what is now known as *quantum mechanics*.

provement over the Bohr–Sommerfeld (or classical) treatment in other respects. However, these advantages are somewhat offset by a loss in terms of physical meaning. For one thing, we can no longer think of an electron as moving in a definite orbit. The stationary orbit of an electron of given angular momentum is replaced by a mathematical expression, sometimes referred to as an *orbital*.

In accordance with what was said earlier, we may think of the orbital as determining the statistical probability of finding the specified electron at any position within the atom. Sometimes the orbital is represented as a cloud of varying density. This must not be taken to mean that the electron is smeared out into a cloud of varying density; rather, the density of the cloud at any point gives the probability of finding the electron at that point.

We may now ask: What do the electron orbitals look like? We can obtain some idea by considering the simple example of the hydrogen atom. The probability of finding an electron at any distance r from a hydrogen nucleus, *in any direction*, as calculated from wave mechanics, is shown in the figure on page 70 for three cases in each of which $l = n - 1$: $n = 1$, $l = 0$; $n = 2$, $l = 1$; and $n = 3$, $l = 2$. These correspond to circular orbits of the Bohr–Sommerfeld theory. We can see that there is a small probability that the electron may appear quite close to, or at a considerable distance from, the nucleus. It is of interest, however, that the maximum probability of finding the electron occurs at distances equal to the radii of Bohr's stationary circular orbits.

A somewhat similar situation arises for other combinations of the quantum numbers n and l , equivalent to Bohr–Sommerfeld elliptical orbits. The probability (as a function of distance from a hydrogen nucleus) for an electron with $n = 3$ and $l = 1$ is depicted in the figure on page 71. There are two locations where the electron is most likely to be found; one is near the nucleus and the other farther away, the latter being the more probable. This corresponds to the



The probability of finding electrons at a specified distance r from a nucleus in any direction, the distances are given in units of 10^{-8} cm. When $l = n - 1$, there is a single maximum at a distance equal to the Bohr stationary orbit radius, it is at this distance that the probability of finding the electron is greater than at any other distance

physical interpretation of an elliptical orbit given on p. 56. However, the analogy cannot be extended too far; for $n = 3$, $l = 0$, there are three maxima in the probability curve. The third one, which is much smaller than the others, is quite close to the nucleus.

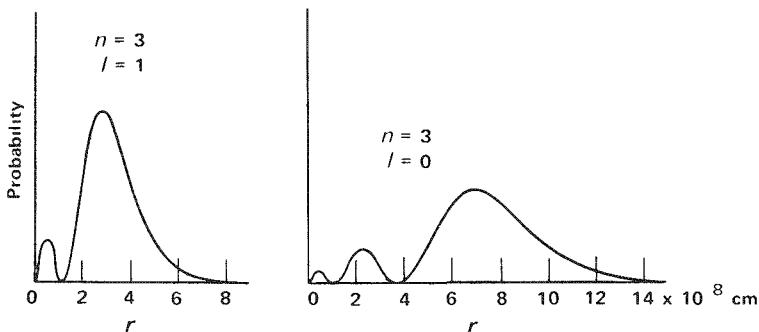
To sum up, we may still think of electrons revolving about the atomic nucleus, since an electron has a particular angular momentum determined by the quantum numbers l and s . But this revolution does not take place in definite orbits. The distance of an electron from the nucleus at any instant cannot be known, although the electron spends most of its time in the region where a Bohr-Sommerfeld orbit would be.

It is still the practice in some publications to picture an atom as a central nucleus with a number of electrons revolving around it in definite orbits, oriented in various directions in space. Such pictures are misleading, and they should not be interpreted literally. For one thing, there are no definite orbits; and for another, their equivalents, the orbitals, are in three dimensions, whereas an orbit in these pictures has two dimensions only.

The Arrangement of Electrons in Atoms

We have now laid the groundwork for the existence of the electron quantum numbers n , l , m , and s , and we are in a position to go on to say something about the arrangement (or configuration) of the electrons in an atom. Certain restrictions were given earlier relating to the permissible values of l for a given n , of m for each l , and of s .* Before we can say

*Remember $l = n - 1, n - 2, n - 3, \dots 0, m = l, l - 1, l - 2, \dots, 0, -(l - 2), -(l - 1), -l, s = +\frac{1}{2}, -\frac{1}{2}$.



Probability of finding $n = 3, l = 1$ and $n = 3, l = 0$ electrons at various distances from a nucleus, the probability curve for $n = 3, l = 2$ is given at the right of the preceding figure.

anything about the quantum numbers of the electrons in an atom, there must be a further condition. This was provided in 1925 by the Austrian mathematical physicist Wolfgang Pauli who introduced an important general rule, called the *exclusion principle*. As applied to the present situation, it may be stated in the following form: No two electrons in an atom can have the same four quantum numbers.

Let us see what this means. Consider, first, an electron for which $n = 1$; the only possible value of l is then 0, and m must also be zero. Finally, the value of s can be either $+\frac{1}{2}$ or $-\frac{1}{2}$. It is clear that there can be only two electrons for which n is 1; their four quantum numbers are indicated below.

n	1	1
l	0	0
m	0	0
s	$+\frac{1}{2}$	$-\frac{1}{2}$

There is no other possibility that will satisfy the restrictions on l , m , and s for $n = 1$, and not violate the exclusion principle. Hence, only two electrons can occupy the first group, or shell, for which n is 1.

We may turn now to the second shell of electrons for which $n = 2$. The permissible values of l are 1 and 0; if $l = 1$, then m can be 1, 0, -1 , whereas for $l = 0$, m is also 0. In each case, s may be either $+\frac{1}{2}$ or $-\frac{1}{2}$. There are now eight possibilities satisfying the requirements for $n = 2$; thus,

n	2	2	2	2	2	2	2	2
l	0	0	1	1	1	1	1	1
m	0	0	1	1	0	0	-1	-1
s	$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$	$+\frac{1}{2}$	$-\frac{1}{2}$

Consequently eight electrons can occupy the second shell.

By working in this manner, the results in the accompanying table can be derived for the maximum



Wolfgang Pauli

numbers of electrons in the first five shells, which have $n = 1, 2, 3, 4, 5$, respectively. It is seen that the successive completed shells can contain 2, 8, 18, 32,

Maximum Numbers of Orbital Electrons in Shells

n	$l = 0$	1	2	3	4	Total
1	2	-	-	-	-	2
2	2	6	-	-	-	8
3	2	6	10	-	-	18
4	2	6	10	14	-	32
5	2	6	10	14	18	50

and 50 electrons. These numbers may be written as 2×1^2 , 2×2^2 , 2×3^2 , 2×4^2 , and 2×5^2 , respectively. This is in accord with the general idea, derived from both the Bohr theory and from wave mechanics, that the larger the principal quantum number n is, the more likely are electrons to be found at greater distances from the nucleus. The greater the distance, the more the space available in the atom, and hence the larger the number of electrons that can be accommodated.

The table gives the maximum number of electrons in each principal shell as well as in the subshells, that is, for the different l values for each value of n . By using these numbers in conjunction with information obtained from atomic spectra, from the characteristic X rays, and from the chemical properties of the elements, it has been possible to deduce the arrangement of the orbital electrons in the atoms of most elements. As a simple example we may consider the element sodium; this has an atomic number of 11, and so there are 11 orbital electrons in the atom. These are arranged with two electrons in the first shell ($n = 1$), eight in the second shell ($n = 2$), and one in the third shell ($n = 3$). Thus, the first two electron shells are filled and there is one electron over, which must go into the third shell. Both the atomic spectrum and

the chemical properties are consistent with there being a single electron outside a set of closed shells.

Many interesting results have been obtained from a study of electronic configurations that cannot be discussed here. We will mention, however, that the repetition observed in the periodic table of the chemical and some of the physical properties of the elements can be readily accounted for. It will be noted that the numbers of elements in successive periods given in the table in the figure on page 28 are 2, 8, 8, 18, 18, 32; these are clearly related to the maximum numbers of electrons in successive shells and subshells.





In this century the universe has been extended outwardly from our own galaxy, the star group of our Milky Way, to enormously greater reaches of space and inwardly from the periphery of the atom to entities a million times smaller. The wonder is that the human mind can conceive at all of these realms so remote from the environment in which it had its slow development over a period of a thousand centuries.

Victor Guillemin

THE NUCLEUS

Nuclear Binding Energy

In the concluding section of this booklet, we will say something more about the nature of the atomic nucleus. We saw earlier (p. 41) that the mass of a nucleus is always less than the sum of the masses of the individual Z protons and $A - Z$ neutrons that it contains. What does this mean? It means that the nucleons (the protons and neutrons in the nucleus) are bound together, the energy of binding being directly related to the mass difference. The *binding energy* of a nucleus is defined as the energy which would be released if Z protons and $A - Z$ neutrons could be combined together to form a nucleus. It is equal to the energy that would have to be supplied in the reverse process of breaking up a nucleus into its constituent nucleons.

In order to calculate the binding energy from the mass difference between the nucleus on the one hand, and of the component nucleons on the other hand, we make use of the Einstein mass-energy equation (7). Instead of giving energy in ergs and masses in grams, different units of energy and mass are employed. Masses are conveniently expressed in terms of the atomic mass unit (amu) defined on page 41. For stating energies a unit called the *electron volt* is generally used; it is the energy acquired by a unit electronic charge when it is accelerated by a potential of 1 volt. In studies of nuclear binding energies and nuclear energies in general, millions of electron volts are involved, and so the *million electron volt* unit, abbreviated to MeV, is commonly employed. By using appropriate conversion factors and the known velocity of light in a vacuum, it is found from equation (7) that

$$E \text{ (MeV)} = 931 M \text{ (amu)}$$

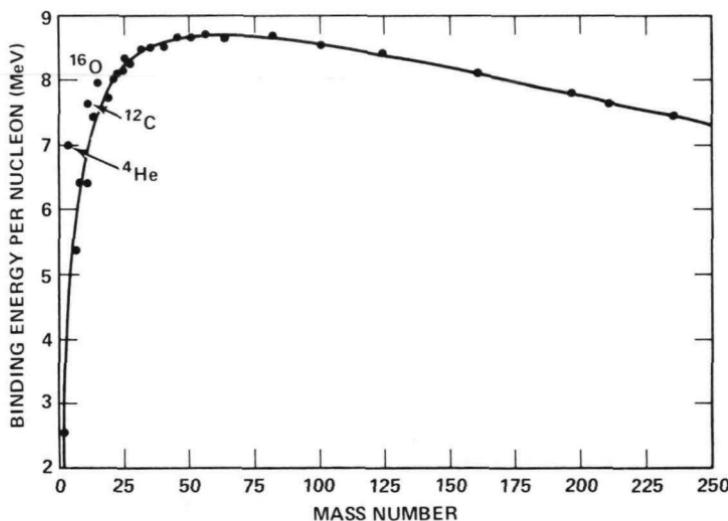
In other words, to find the energy in MeV equivalent to a mass of M amu, the latter is multiplied by 931.

Suppose M_p amu represents the mass of a proton and M_n amu is the mass of a neutron; the total mass of the individual nucleons in a nucleus would then be $ZM_p + (A - Z)M_n$. If M amu is the actual mass of the nucleus, the mass difference, equivalent to the binding energy (B.E.) of the nucleus, is represented by

$$\text{Mass difference in amu} = [ZM_p + (A - Z)M_n] - M$$

The B.E. in MeV is then obtained upon multiplying this mass difference by 931. For actual determination of the mass difference, the masses of the Z orbital electrons in the atom are included; ZM_p then becomes ZM_H , where M_H is the mass of the hydrogen atom, a proton plus an electron, and M is the atomic mass rather than the nuclear mass.

Since M_H and M_n are known and M can be determined from mass spectrometer measurements, all the information is



The variation with mass number of the average binding energy per nucleon in several stable nuclei. For all except the lightest nuclei, the average binding energy per nucleon is approximately 8 MeV.

available for calculating the B.E. of any nucleus. The total B.E. increases steadily with the number of nucleons in the nucleus; and so, for purposes of graphical representation, it is convenient to use the average B.E. per nucleon. This is obtained by dividing the total B.E. by the mass number A , which is equal to the number of nucleons in a given nucleus.

The variation with mass number of the average B.E. per nucleon, obtained as described above, for several stable nuclei is shown in the figure. It is clear that most of the values lie on, or close to, a continuous curve; this curve has several interesting features. For example, we can see that with increasing mass number the average B.E. per nucleon increases to a broad maximum, after which it decreases slowly. This tells us immediately what kinds of nuclear reactions can lead to the release of energy.

The Release of Nuclear Energy

Any nuclear process accompanied by an increase in the average B.E. per nucleon must be associated with the liberation of energy. Fundamentally, the reason is that in such a process there is a decrease in mass, and there must consequently be a release of an equivalent amount of energy. It is evident from the figure that there are two general types of nuclear reactions that can lead to the production of energy. One is the combination or interaction of two of the lighter nuclei to form a heavier one; this is called *nuclear fusion*. The other is the division or splitting of a heavy nucleus to form two (or more) lighter ones; this process is known as *nuclear fission*.

The fusion of the nuclei of the lightest element, hydrogen, to form a heavier nucleus, that of helium-4, is the source of energy of the sun and other similar stars. In some of the hotter stars, helium-4 nuclei combine to produce nuclei of carbon-12 and oxygen-16. Astrophysicists think that nuclei as heavy as iron-56 are formed in stars by nuclear fusion, with an accompanying release of energy. Beyond the

mass number 56 the release of energy by fusion is not possible because the average B.E. per nucleon decreases, as shown in the figure.

Fusion of the isotopes of hydrogen, with the release of large amounts of energy, takes place in the hydrogen (thermonuclear) bomb. Efforts are being made to find ways in which this fusion energy can be used as a controlled source of power for peaceful purposes.*

The release of energy by the fission of heavy nuclei, such as those of the uranium-235 isotope, has been a practical reality for over a quarter of a century. Methods have been developed for making this energy available for the production of electric power in a device called a nuclear reactor.† The energy of fission has also been used in the so-called atomic bomb.

Another point to notice in the figure is that, except for nuclei of very small mass number, the average B.E. per nucleon is always in the range of about 7.5 to 8.5 MeV. That is to say, on the average, the binding energy of a nucleon (a single neutron or proton) in a nucleus is around 8 MeV. Of course, the individual binding energies vary somewhat, say from about 6 to 10 MeV, but the average is in the vicinity of 8 MeV per nucleon for the great majority of nuclei. The significance of this quantity will be explained shortly.

Nucleon and Electron Binding Energies

For purposes of comparison, let us consider the binding energy of an orbital electron in an atom. This energy, which can either be calculated or in many instances measured, varies with the energy level of the electron, i.e., with its average

*See *Controlled Nuclear Fusion*, another booklet in this series.

†For descriptions of various types of nuclear reactors, see other booklets in this series, such as *Nuclear Reactors* and *Nuclear Power Plants*

distance from a given nucleus. It has been found, however, that electron binding energies are usually not more than 10 electron volts. Thus the average binding energy of a nucleon in an atomic nucleus, roughly 8 MeV, is about a million times greater than that of an orbital electron. The forces which hold the particles together inside a nucleus are thus very much stronger than the electrical forces that bind the electrons to the nucleus.

One of the consequences of the very strong nuclear force is that the energy released in such nuclear reactions as fusion and fission is very much more "concentrated" than other more familiar forms of energy. The chief common (or conventional) sources of energy—such as are used, for example, in the production of electricity—are coal, petroleum oil, and natural gas. These conventional fuels release energy when they undergo chemical reaction with oxygen in the air. This reaction, referred to as combustion, involves a rearrangement of the electrons in atoms.

We can see, therefore, that *nuclear energy*, sometimes less precisely called atomic energy, results from the rearrangement of nucleons in nuclei, whereas the more conventional forms of energy arise from electronic rearrangements. In view of the million-fold difference in the binding energies, it is easy to understand why nuclear energy is so much more "concentrated". For example, about the same amount of energy is produced by the complete fission of 1 pound of uranium-235 as by the combustion of about 3,000,000 pounds of coal.

Nuclear Forces

What do we know about the forces that hold a nucleus together? The protons all carry a positive charge of electricity and so they must repel one another. Yet the nucleons, both protons and neutrons, are tightly bound in the nucleus. Obviously, there must also be a very strong force of attraction. This force cannot be electrical in nature since such

a force would lead to repulsion. Furthermore, an electrical force could not involve the neutrons since they have no electrical charge. There is evidently an entirely different type of force operating in a nucleus, and it must be so strong that it can more than overcome the electrical repulsion of the closely packed protons. This special nuclear force is an aspect of what is called the *strong interaction*; it is the strongest type of force known.*

Although there are many things about this strong nuclear force that we do *not* know, some of its aspects are fairly obvious. First, it must have a very short range because it is effective only over very short distances. Thus, it is extremely strong over a distance of about a nuclear radius (10^{-13} to 10^{-12} cm); beyond this distance, the force drops sharply almost to zero. It differs in this respect from the more familiar electrical and gravitational forces; these two forces can operate over relatively long ranges and they fall off gradually with increasing distance.

Another thing we are sure of is that the nuclear force of attraction exists between protons and protons, as well as between neutrons and neutrons, and neutrons and protons. These forces are referred to as *p-p*, *n-n*, and *n-p* forces, respectively, the symbol *p* standing for a proton and *n* for a neutron.

By analyzing the binding energies of various nuclei, containing different numbers of protons and neutrons, the conclusion has been reached that the *p-p*, *n-n*, and *n-p* nuclear forces are almost equal in strength. Clearly, the strong nuclear force has nothing to do with electrical charge.

*For more about the strong interaction (and other types of forces), see the companion booklet *Microstructure of Matter*

Origin of Nuclear Forces

The problem of understanding the nuclear force is an extremely difficult one. The only approach that seems at all reasonable is based on an idea put forward by Hideki Yukawa in Japan in 1935. At the present time, it is thought that the nuclear force arises from the emission and absorption of a particle called a *pi-meson* or *pion*. This particle, which is known to exist in positively charged, negatively charged, and neutral forms, has a mass about 270 times that of an electron, or about one-seventh the mass of a proton or a neutron.*

Many scientists think that the attraction between two nucleons, leading to the *p-p*, *n-n*, and *n-p* forces, arises from the transfer of a pion from one nucleon to another. It is supposed that protons and neutrons are continually emitting and absorbing pions, either electrically charged or neutral pions. The particles exchanged, which remain free for such an extremely short time (about 4×10^{-24} second) that they are undetectable, are referred to as *virtual pions*. Calculations show that if the pion moves with the velocity of light, it can travel a distance from the nucleon emitting it of no more than 1.5×10^{-13} cm during the short time it remains free.

A nucleon is thought of as being surrounded by a "cloud" of these virtual pions. When two nucleons get closer than 1.5×10^{-13} cm, a pion can pass from one to the other and a force of attraction is exerted between them. The pion does not remain free long enough to traverse a greater distance, and so the force could not operate over more than about 1.5×10^{-13} cm. This would account for the short-range character of the nuclear force. The attraction between any two nucleons, regardless of whether they are electrically charged (protons) or not (neutrons), could arise from the transfer of either a charged or an uncharged pion, according to the circumstances.

*For more about pions, see *Microstructure of Matter* in this series.

Nuclear Models

We saw earlier that much is known about the arrangement of the orbital electrons in an atom. This is largely because the force between the electrons and the nucleus, as well as between the electrons themselves, is electrical in nature and the mathematical treatment of such forces is well established. Unfortunately, there is no satisfactory mathematics for treating the strong nuclear interaction based on the concept of pion transfer outlined above. Consequently, there is no simple theory of the detailed structure of the nucleus.

What physicists have done so far is to propose "models" which can be used to interpret certain aspects of the behavior of nuclei. Ultimately, no doubt, there will be a single nuclear model that accounts for all the properties, just as there is now a single model for the orbital electrons. A few of the more important nuclear models will be described briefly here.

The first useful model of the nucleus, called the *liquid-drop model*, was proposed in 1936 by Bohr. Some of the successful applications of the model are in the semi-theoretical calculation of nuclear binding energies, in the study of fission, and in the prediction of the relative stabilities of nuclei of equal mass number. It has also been used in the theory of the *compound nucleus*; such a nucleus is supposed to be formed when a nucleus captures a particle, especially a neutron, of not too high energy.

An ordinary liquid is composed of molecules (p. 8) that are packed fairly closely together. Each molecule in a liquid attracts those with which it is in contact, but has little or no influence on more distant molecules. Much the same is true in a nucleus because of the very short range of the forces between nucleons.

A liquid drop exhibits the phenomenon called *surface tension*, as a result of which the surface of a drop tends to contract so as to have the smallest possible area. There are

reasons for believing that a nucleus behaves in a similar manner.

Furthermore, in a drop of liquid, both the total internal energy and the volume are proportional to the number of molecules present in the drop. Atomic nuclei have similar characteristics: The binding energy and the volume are roughly proportional to the number of nucleons present in the nucleus. How do we know this? We saw that the average binding energy per nucleon is almost constant, about 8 MeV, for nearly all the different nuclei. In other words, the total binding energy of a nucleus is proportional to the number of nucleons, that is, to the mass number A . We also saw (p. 25) that the radius of a nucleus is proportional to $\sqrt[3]{A}$. Hence, the volume, which depends on the cube of the radius for a spherical nucleus, is also proportional to A , the number of nucleons.

Another important nuclear model is the *shell model*, sometimes also known as the *independent particle model*. The possibility that the nucleons inside a nucleus might form a series of shells, just as the orbital electrons do, was first considered in 1932; but this model received little attention for several years. In 1948, however, Maria Goeppert Mayer in the United States brought together a considerable body of evidence showing that nuclei containing either 2, 8, 20, 50, or 82 protons or 2, 8, 20, 50, 82, or 126 neutrons (or both) were exceptionally stable. These numbers of protons or neutrons, which were called *magic numbers*, would presumably represent successive completed (closed) shells of 2, 6, 12, 30, 32, and 44 nucleons, analogous to the maximum numbers of orbital electrons (p. 74). These numbers of nucleons are obtained by subtracting each magic number from the next one starting with zero.

The first attempts to account for these numbers of nucleons in completed shells were unsuccessful. But in 1949, Mayer in the United States and Otto Haxel, J. Hans D. Jensen, and Hans E. Suess in Germany, proposed essen-



Maria Goeppert Mayer

Maria Goeppert Mayer

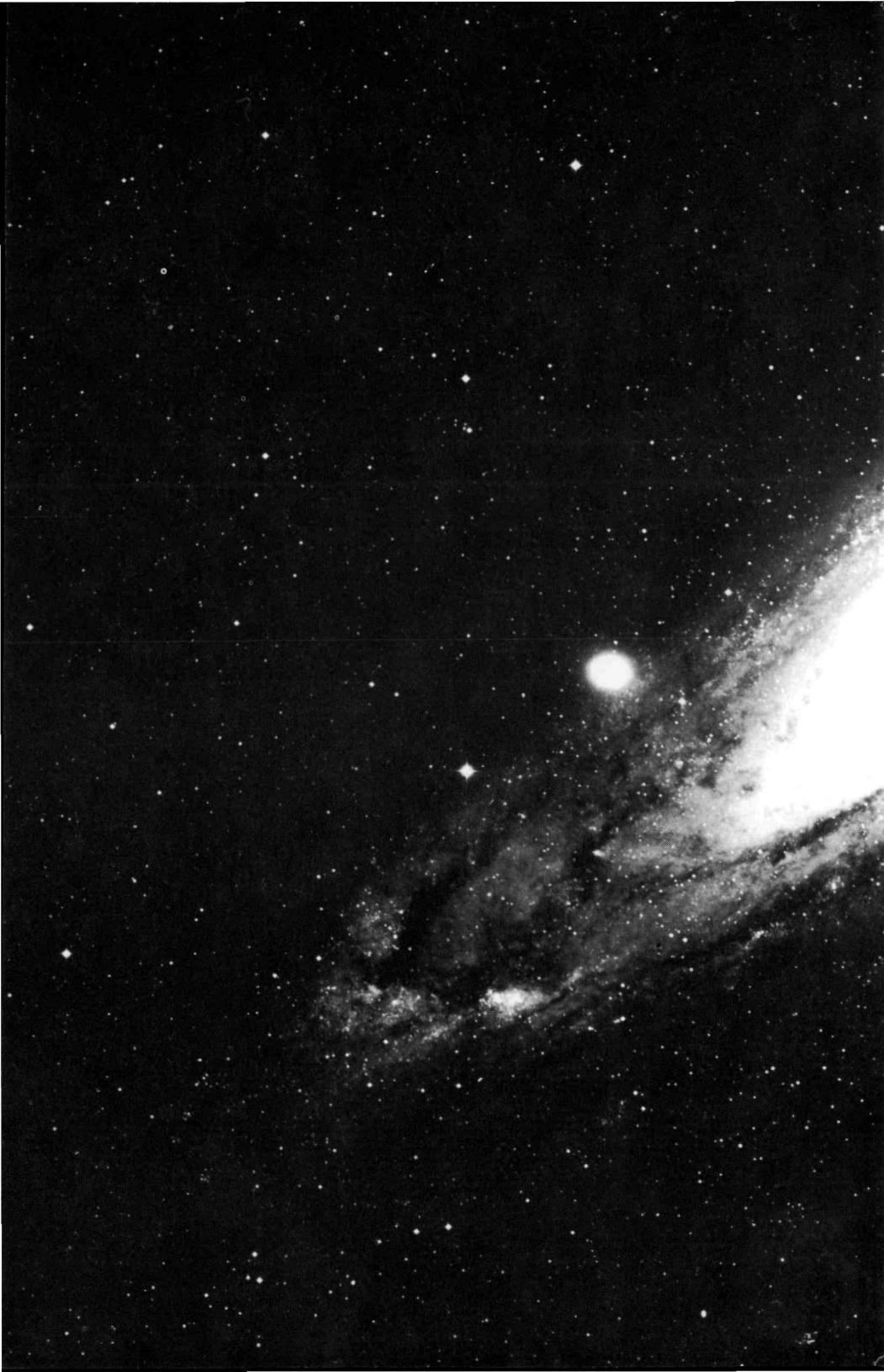
tially the same solution to the problem. The conditions in a nucleus must be different from those applicable to electrons because the numbers of particles in closed shells are generally not the same for nucleons and electrons. Since the same (or similar) four quantum numbers must be applicable to nucleons, as well as to electrons, where does this difference lie?

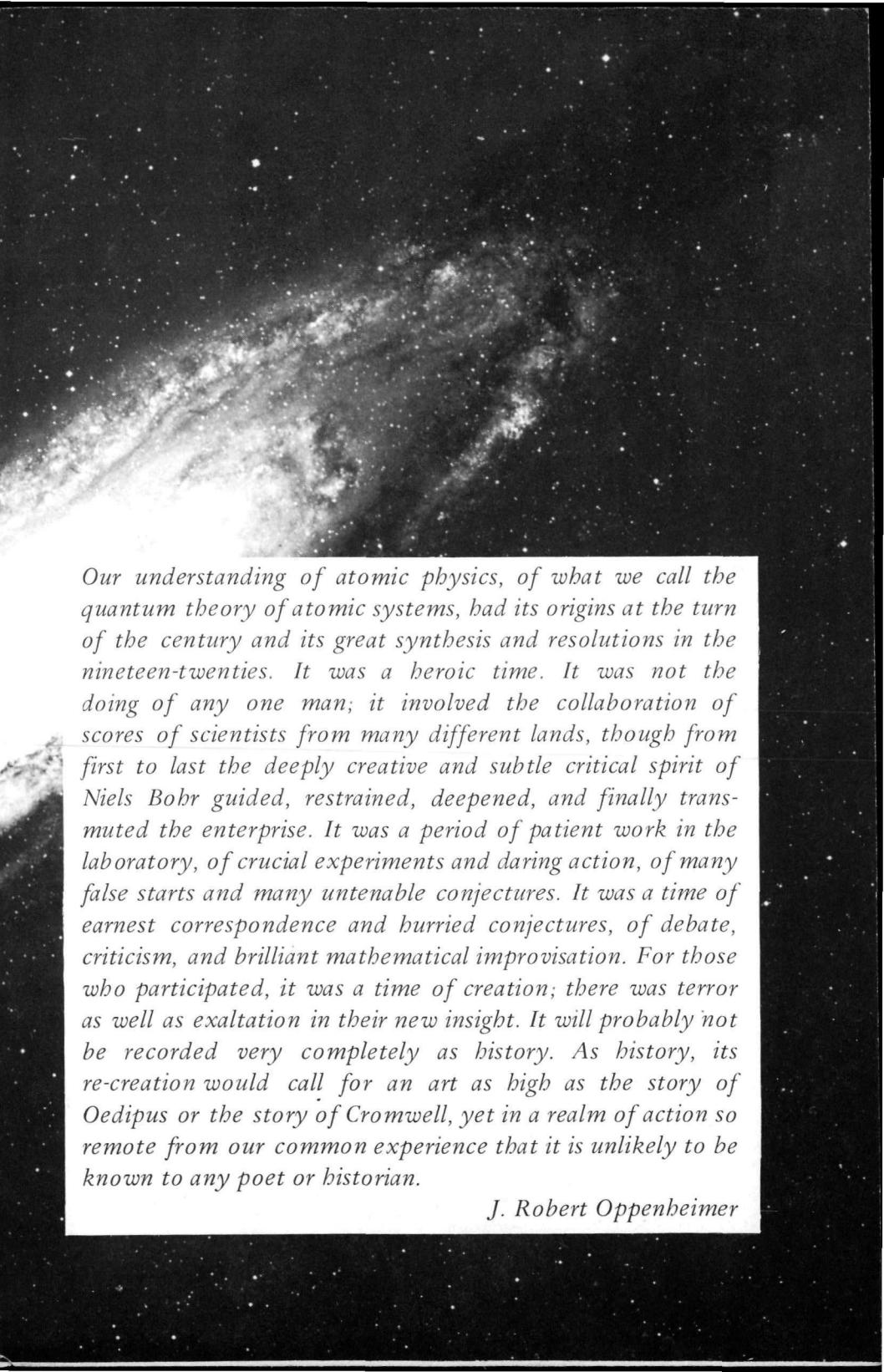
From the behavior of the orbital electrons, scientists have concluded that, as a general rule, there is no interaction between the orbital angular momentum and the spin angular momentum of any single electron. In the nucleus, however, such interaction apparently does occur between the orbital and spin angular momenta of each individual nucleon. By assuming this to be the case and making use of the Pauli exclusion principle (p. 72), which also applies to nucleons, the magic numbers can be derived from purely theoretical considerations.

Many properties of nuclei have been explained in terms of the arrangement in successive shells of the constituent protons and nucleons. For example, nuclear spins of different nuclei in various energy states have been accounted for successfully. In some respects, however, the shell model proved to be unsatisfactory. Thus, in the original form of the model, the core of closed shells in a nucleus was supposed to be spherical even when there were nucleons outside the closed shells.

In the *collective model* of the nucleus, the central core is assumed to behave like a liquid drop that can suffer deformation in certain circumstances. If this deformation is not stationary, but travels around the nucleus, the resulting motions are equivalent to oscillations and rotations of the nucleus as a whole. By use of this idea, in conjunction with essentially the same mathematics as is used to study the oscillations and rotations of molecules, it has been possible to account for the known energy levels (or energy states) of certain nuclei.

The collective model is, in a sense, a modification of the liquid drop model and does not utilize the details of the shell model. It was natural, therefore, for physicists to attempt to combine the valuable features of each model. This has been done in the *unified model* of the nucleus, which successfully reconciles many different aspects of nuclear behavior. Furthermore, the unified model provides justification for using the simple liquid-drop model in connection with fission and with compound nucleus formation.





Our understanding of atomic physics, of what we call the quantum theory of atomic systems, had its origins at the turn of the century and its great synthesis and resolutions in the nineteen-twenties. It was a heroic time. It was not the doing of any one man; it involved the collaboration of scores of scientists from many different lands, though from first to last the deeply creative and subtle critical spirit of Niels Bohr guided, restrained, deepened, and finally transmuted the enterprise. It was a period of patient work in the laboratory, of crucial experiments and daring action, of many false starts and many untenable conjectures. It was a time of earnest correspondence and hurried conjectures, of debate, criticism, and brilliant mathematical improvisation. For those who participated, it was a time of creation; there was terror as well as exaltation in their new insight. It will probably not be recorded very completely as history. As history, its re-creation would call for an art as high as the story of Oedipus or the story of Cromwell, yet in a realm of action so remote from our common experience that it is unlikely to be known to any poet or historian.

J. Robert Oppenheimer

CONCLUSION

We can see from the information presented in this booklet that much is known about atomic structure. Much of the behavior of electrons is well understood—even some behavior that is too complicated to describe here—and the forces between electrons and nuclei, as well as between electrons themselves, can be calculated quite accurately. The same is not true, however, for the particles in the nucleus. What is lacking is an adequate way to describe the nuclear force. At the present time it appears that a resolution of this situation may depend either on the development of a new mathematical approach or perhaps on some new and unexpected discovery in nuclear physics. Until nuclear forces can be calculated by purely theoretical mathematical methods, the story of the structure of the atom must be regarded as unfinished.

Quotations

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Basic Books

Atomic Energy: The Story of Nuclear Science, Irene D. Jaworski and Alexander Joseph, Harcourt Brace Jovanovich, Inc., New York, 1961, 218 pp., \$4.95. Grades 8–12. Describes the nature and structure of the atom and presents many safe home experiments, such as producing and controlling an electron beam, photographing alpha tracks, observing scintillations, making a reactor model, and constructing a Van de Graaff generator.

Biography of Physics, George Gamow, Harper and Row, Publishers, New York, 1961, 338 pp., \$6.50 (hardback); \$2.45 (paperback). Grades 9–12. A history of theoretical physics.

The Heart of the Atom: The Structure of the Atomic Nucleus, Bernard L. Cohen, Doubleday and Company, Inc., New York, 1967, 107 pp., \$4.50 (hardback); \$1.25 (paperback). Grades 9–12. Describes all aspects of this “atomic heart”: its structure, motion, radiation, and large-scale application.

Inside the Atom (revised edition), Isaac Asimov, Abelard–Schuman, Ltd., New York, 1966, 197 pp., \$4.95. Grades 7–10. This comprehensive, well-written text explains nuclear energy and its applications.

Nuclear Physics and the Fundamental Particles, H. H. Heckman and P. W. Starring, Holt, Rinehart and Winston, Inc., New York, 1963, 410 pp. (Out of print.) Grades 10–12. A nonmathematical textbook written for high school students. The basic science of the nucleus is stressed.

The Questioners: Physicists and the Quantum Theory, Barbara Lovett Cline, Crowell Collier and Macmillan, Inc., New York, 1965, 274 pp., \$5.00 (hardback). Available in paperback with the title *Men Who Made a New Physics: Physicists and the Quantum Theory*, New American Library, Inc., New York, \$0.75. Grades 9–12. An excep-

tionally well-delineated and personable account of the development of the quantum theory by physicists in the first quarter of this century.

The Restless Atom, Alfred Romer, Doubleday and Company, Inc., New York, 1960, 198 pp., \$1.25. Grades 9-12. A stimulating nonmathematical account of the classic early experiments that advanced knowledge about atomic particles.

Rutherford and the Nature of the Atom, Edward N. da C. Andrade, Doubleday and Company, Inc., New York, 1964, 218 pp., \$1.25. Grades 10-12. Andrade was one of Rutherford's assistants at the University of Manchester.

Secrets of the Nucleus, Joseph S. Levinger, National Science Teachers Association, Washington, D. C., 1967, 127 pp., \$0.50. Grades 9-12. This introduction to nuclear energy includes science projects and experiments.

Advanced Books

The Atom and Its Nucleus, George Gamow, Prentice-Hall, Inc., New York, 1961, 153 pp., \$1.95. A popular-level discussion of nuclear structure and the applications of nuclear energy.

The Atomists (1805-1933), Basil Schonland, Oxford University Press, Inc., New York, 1968, 198 pp., \$5.60. This book, which can be understood by anyone who has had a high school physics course, presents atomic theory development from Dalton through Bohr. It achieves a good balance between popular treatments and highly technical works without slighting the technical aspects.

The Born-Einstein Letters: The Correspondence Between Albert Einstein and Max and Hedwig Born from 1916 to 1955, commentaries by Max Born, translated by Irene Born, Walker and Company, New York, 1971, 240 pp.,

\$8.50. These interesting letters reveal the scientific and personal lives of two famous scientists.

The Discovery of the Electron: The Development of the Atomic Concept of Electricity, David L. Anderson, Van Nostrand Reinhold Company, New York, 1964, 138 pp., \$1.50. This semitechnical book describes the experimental and conceptual developments that led to the discovery of the electron.

The Electron: Its Isolation and Measurement and the Determination of Some of Its Properties, Robert Andrews Millikan, 1917 edition edited by Jesse W. M. DuMond and reissued in 1963, 268 pp., University of Chicago Press, Chicago, Illinois, \$6.00 (hardback); \$3.95 (paperback). These researches won for Millikan the Nobel Prize for Physics in 1923. An introduction by an associate of the author puts the discoveries in perspective.

Elementary Particles, David H. Frisch and Alan M. Thordike, Van Nostrand Reinhold Company, New York, 1964, 153 pp., \$1.75. An account of the basic properties of particles and the experimental techniques used to study them.

Elementary Particles: A Short History of Some Discoveries in Atomic Physics, Chen Ning Yang, Princeton University Press, Princeton, New Jersey, 1961, 65 pp., \$4.50. Dr. Yang was a co-winner of the Nobel Prize along with Dr. Tsung-Dao Lee for suggesting the experiments that led to the downfall of the conservation of parity principle. Here he provides a general outline for laymen of the history of elementary particle research during the last 60 years.

The Evolution of Physics: The Growth of Ideas from Early Concepts to Relativity and Quanta, Albert Einstein and Leopold Infeld, Simon and Schuster, Inc., New York, 1938, 302 pp., \$6.95 (hardback); \$1.95 (paperback). Traces the steps from the mechanical view of the universe held by the classical physicists through subsequent developments that led to quantum mechanics.

Inside the Nucleus, Irving Adler, The John Day Company, Inc., New York, 1963, 192 pp., \$4.95. An explanation of the structure of the atom and the amazing discoveries in recent years about its nucleus.

J. J. Thomson: Discoverer of the Electron, George Thomson, Doubleday and Company, Inc., New York, 1966, 240 pp., \$1.45. This biography, written by J. J. Thomson's son, describes his research at the famed Cavendish Laboratory in Cambridge, England.

John Dalton and the Atomic Theory: The Biography of a Natural Philosopher, Elizabeth C. Patterson, Doubleday and Company, Inc., New York, 1970, 320 pp., \$5.95 (hardback); \$1.95 (paperback). The drama of Dalton's life—his rigorous self-teaching, scientific work, and struggle to overcome class barriers in 19th century England—are well presented. Quotations from letters, diaries, and published works give a clear picture of Dalton's atomic theory research and his time.

Matter and Light: The New Physics, Louis Victor de Broglie, translated by W. H. Johnston, Dover Publications, Inc., New York, 1939, 300 pp., \$2.50. These essays on physics, which include the author's Nobel Prize speech, were written by one of the pioneers in quantum mechanics.

The Neutrino: Ghost Particle of the Atom, Isaac Asimov, Doubleday and Company, Inc., New York, 1966, 223 pp., \$5.50 (hardback); \$1.95 (paperback) from Dell Publishing Company, Inc., New York. The author traces a century-long chain of events that proved, to the surprise of scientists, that a strange little particle called the neutrino actually exists and is produced in astronomical numbers inside the sun and other stars.

Niels Bohr: The Man, His Science, and the World They Changed, Ruth Moore, Alfred Knopf, Inc., New York, 1966, 436 pp., \$7.95. An interesting biography of one of the pioneers in the study of the internal structure of the atom.

Physics and Beyond: Encounters and Conversations, Werner Heisenberg, translated by Arthur J. Pomerans, Harper and Row, Publishers, New York, 1970, 247 pp., \$7.95. Werner Heisenberg, a Nobel Prize physicist, presents his autobiography in the form of conversations with such men as Max Planck, Albert Einstein, Niels Bohr, Ernest Rutherford, Otto Hahn, and Enrico Fermi.

Physics for Poets, Robert H. March, McGraw-Hill Book Company, New York, 1970, 302 pp., \$7.50. A physics textbook for nonscience students. The book covers certain developments of classical mechanics, relativity, and atomic and quantum physics. With this book the author won the 1971 American Institute of Physics—U. S. Steel Foundation Science Writing Award in Physics and Astronomy.

Sourcebook on Atomic Energy (third edition), Samuel Glasstone, Van Nostrand Reinhold Company, New York, 1967, 883 pp., \$15.00. A standard reference work, written for both scientists and the public.

The Story of Quantum Mechanics, Victor Guillemin, Charles Scribners Sons, New York, 1968, 332 pp., \$9.95 (hardback); \$2.95 (paperback). A complex subject is presented in a clear and fascinating way in this beautifully written book. Philosophical as well as scientific implications of quantum mechanics are discussed. A glossary and a well-annotated reference list are included.

Thirty Years That Shook Physics: The Story of Quantum Theory, George Gamow, Doubleday and Company, Inc., New York, 1966, 224 pp., \$5.95 (hardback); \$1.45 (paperback). The development of the quantum theory presented in nontechnical language.

The World of the Atom, 2 volumes, Henry A. Boorse and Lloyd Motz (Eds.), Basic Books, Inc., Publishers, New York, 1966, 1873 pp., \$35.00. Contains the actual text of landmark documents in the history of atomic physics, each preceded by commentary that places it in the context of the discoverer's personal life and in the conditions prevailing in science and in society in his time.

Motion Pictures

Available for loan without charge from the AEC Headquarters Film Library, Office of Information Services, U. S. Atomic Energy Commission, Washington, D. C. 20545.

Atomic Physics, 90 minutes, black and white, 1948.

Produced by J Arthur Rank Organisation, Ltd. The history of atomic energy beginning with Dalton's theory is described. Some of the highlights include: Faraday's early experiments in electrolysis, Mendeléev's periodic table, investigation of cathode rays, discovery of the electron, discovery of X rays, discovery of the neutron, and uranium fission. One sees and hears such famous scientists as J. M. Thomson, Ernest Rutherford, Albert Einstein, J. D. Cockcroft, and O. R. Frisch.

Exploring the Atomic Nucleus, 13½ minutes, color, 1969.

Produced by Coronet Films. Particle accelerators, the basic tools of high energy physics used to explore the atomic nucleus, are described. The purpose of the film is to show some of the recent discoveries physicists have made concerning nuclear structure, the basic equipment used, and how the resulting data are analyzed.

The Mass of Atoms, 47 minutes, black and white, 1966.

Produced by Educational Services, Inc. for the College Physics Film Program under a grant from the National Science Foundation. The step-by-step operations in an experiment performed by two scientists to determine the masses of a helium atom and a polonium atom are described. Various laboratory techniques and precautions necessary for these measurements are shown in detail.

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